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A PROBLEM COURSE IN CHEMISTRY

BY

DAVID B. PUGH AND D. EDWIN MILLER

SCHENLEY HIGH SCHOOL, PITTSBURGH

ALLYN AND BACON
BOSTON NEW YORK CHICAGO
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PREFACE

This course uses chemical operations to teach the pupil to think in terms of chemistry. The numerous problems offer little mathematical difficulty, but much drill in chemical science. The mathematical element is merely incidental; the aim is to lead the pupil to understand the chemical principle involved in the problem.

The object has been not to set an arbitrary standard for other teachers to follow, but to compile such a set of problems as will meet the needs of the individual teacher and the individual student.

In the interest of variety and thoroughness, more than six hundred problems are furnished, although the authors have found in their own classes that one hundred and fifty of these exercises, related to the laboratory and text-book material, provide a satisfactorily balanced year's work.

The fundamental principles essential to an intelligent understanding of the calculations are presented in a clear and simple manner. Emphasis is laid upon the fact that all relationships in chemistry are quantitative, and means are provided for the solution of problems commonly met in general chemistry. It has been deemed advisable to call attention to the typical and characteristic reactions of the chemistry of the home, of the laboratory, and of industry. The data provided in the tables have been collected from standard works of reference.

D. B. P.
D. E. M.

Pittsburgh,
March, 1921

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INTRODUCTION

In taking up the study of chemistry and solving problems involving chemical operations, we must know the names and symbols of the various chemical elements, the weights and boiling points of various gases, and certain other matters treated in this Introduction. Above all, we must be thoroughly familiar with the metric system, since this system is practically universal outside of the United States and Great Britain. Even in these countries nearly all scientific calculations and an increasing number of commercial transactions are based upon the metric system.

The need of a general understanding of chemistry is daily becoming more evident. If America is to hold her place among the nations, the engineers and chemists of the future must be trained definitely to apply their knowledge. With the expansion of chemical industry as a result of the late war has come a growing need for men thoroughly trained in ways of obtaining accurate quantitative results.

Chemistry ceased to be merely an art and became a science when Lavoisier introduced the balance into his experiments. In this science the quantitative relationships of materials involved are fundamental. An appreciation of the usefulness of chemistry is not possible if only qualitative descriptions are adhered to. The atom, the symbol, the molecule, the formula, the compound, the equation, all represent definite relative quantities of materials; to impress this idea firmly upon the mind of the student is the aim of this course.

SUGGESTED COURSE

THE problems in this suggested course include essential reactions and calculations in General Chemistry. A group of more difficult problems for students of proved ability is given at the end of this list.

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LIST OF ELEMENTS

Name	Symbol	Atomic Weight		Valence	Specific Gravity
		Exact	Working		
1. Aluminum	**Al**	27.10	27.	3	2.7
2. Antimony	**Sb**	120.20	120.	3 or 5	6.62
3. Argon	**A**	39.88	39.9	—	1.38
4. Arsenic	**As**	74.96	75.	3 or 5	5.73
5. Barium	**Ba**	137.37	137.4	2	3.8
6. Bismuth	**Bi**	208.00	208.	3 or 5	9.78
7. Boron	**B**	11.00	11.	3	2.45
8. Bromine	**Br**	79.92	80.	1	3.12
9. Cadmium	**Cd**	112.40	112.4	2	8.65
10. Caesium	**Cs**	132.81	132.8	1	1.87
11. Calcium	**Ca**	40.07	40.	2	1.54
12. Carbon	**C**	12.00	12.	2 or 4	1.88
13. Cerium	**Ce**	140.25	140.3	4 or 3	6.92
14. Chlorine	**Cl**	35.46	35.5	1	2.49
15. Chromium	**Cr**	52.00	52.	3 or 6	6.92
16. Cobalt	**Co**	58.97	59.	2 or 3	8.72
17. Columbium	**Cb**	93.50	93.5	3 or 5	8.4
18. Copper	**Cu**	63.57	63.6	2 or 1	8.93
19. Dysprosium	**Dy**	162.50	162.5	3	
20. Erbium	**Er**	167.70	167.7	3	4.77
21. Europium	**Eu**	152.00	152.	3	
22. Fluorine	**F**	19.00	19.	1	1.31
23. Gadolinium	**Gd**	157.30	157.3	3	1.31
24. Gallium	**Ga**	69.90	69.9	3	5.94
25. Germanium	**Ge**	72.50	72.5	4	5.5
26. Glucinium	**Gl**	9.10	9.1	2	1.85
27. Gold	**Au**	197.20	197.2	3 or 1	19.32
28. Helium	**He**	4.00	4.	—	.137
29. Hydrogen	**H**	1.008	1.	1	.0695
30. Indium	**In**	114.80	114.8	3	7.12
31. Iodine	**I**	126.92	127.	1	4.94
32. Iridium	**Ir**	193.10	193.1	3	22.42
33. Iron	**Fe**	55.84	56.	2 or 3	7.28
34. Krypton	**Kr**	82.92	82.9	—	2.818
35. Lanthanum	**La**	139.00	139.	3	6.155
36. Lead	**Pb**	207.20	207.	2	11.34
37. Lithium	**Li**	6.94	7.	1	.534
38. Lutecium	**Lu**	17.50	17.5	3	
39. Magnesium	**Mg**	24.32	24.3	2	1.74
40. Manganese	**Mn**	54.93	55.	2, 4, 6, 7	7.42
41. Mercury	**Hg**	200.60	200.	1 or 2	13.595
42. Molybdenum	**Mo**	96.00	96.	3 or 6	9.01

LIST OF ELEMENTS—*Continued*

Name	Symbol	Atomic Weight		Valence	Specific Gravity
		Exact	Working		
43. Neodymium	**Nd**	144.30	144.3	3	6.95
44. Neon	**Ne**	20.20	20.2	—	.674
45. Nickel	**Ni**	58.68	58.7	2 or 3	8.6
46. Nitrogen	**N**	14.01	14.	3 or 5	.967
47. Osmium	**Os**	190.90	190.9	2, 3, 4	22.48
48. Oxygen	**O**	16.00	16.	2	1.1053
49. Palladium	**Pd**	106.70	106.7	2 or 4	12.16
50. Phosphorus	**P**	31.04	31.	3 or 5	1.83
51. Platinum	**Pt**	195.20	195.	2 or 4	21.37
52. Potassium	**K**	39.10	39.	1	.87
53. Praesodymium	**Pr**	140.90	140.9	3	6.48
54. Radium	**Ra**	226.00	226.	2	
55. Rhodium	**Rh**	102.90	102.9	3	12.44
56. Rubidium	**Rb**	85.45	85.5	1	1.52
57. Ruthenium	**Ru**	101.70	101.7	2, 3, 4	12.06
58. Samarium	**Sm**	150.40	150.4	3	7.7
59. Scandium	**Sc**	44.10	44.1	3	
60. Selenium	**Se**	79.20	79.2	2 or 4	4.47
61. Silicon	**Si**	28.30	28.3	4	2.42
62. Silver	**Ag**	107.88	108.	1	10.50
63. Sodium	**Na**	23.00	23.	1	.971
64. Strontium	**Sr**	87.63	87.6	2	2.54
65. Sulfur	**S**	32.06	32.	2, 4, 6	2.046
66. Tantalum	**Ta**	181.50	181.5	5	16.6
67. Tellurium	**Te**	127.50	127.5	4 or 6	6.25
68. Terbium	**Tb**	159.20	159.2	3	
69. Thallium	**Tl**	204.00	204.	1 or 3	11.84
70. Thorium	**Th**	232.40	232.4	4	11.2
71. Thulium	**Tm**	168.50	168.5		
72. Tin	**Sn**	118.70	118.7	2 or 4	5.85
73. Titanium	**Ti**	48.10	48.1	4 or 3	4.5
74. Tungsten	**W**	184.00	184.	6	18.7
75. Uranium	**U**	238.20	238.2	6	18.68
76. Vanadium	**V**	51.00	51.	3 or 5	5.69
77. Xenon	**Xe**	130.20	130.2		4.422
78. Ytterbium	**Yb**	173.50	173.5	3	
79. Yttrium	**Y**	88.70	88.7	3	3.80
80. Zinc	**Zn**	65.37	65.4	2	7.0
81. Zirconium	**Zr**	90.60	90.6	4	4.15

WEIGHT IN GRAMS PER LITER OF VARIOUS GASES UNDER STANDARD CONDITIONS, AND BOILING POINTS UNDER PRESSURE OF 760 MILLIMETERS

NAME	WEIGHT OF ONE LITER IN GRAMS 22.4 Liters Weigh One Gram-Molecular Weight	BOILING POINT, °C.
Acetylene	1.161	−83.8
Air	1.29	—
Ammonia	0.759	−33.5
Argon	1.781	−186.0
Carbon Dioxide	1.964	−78.2
Carbon Monoxide	1.250	−190.0
Chlorine	3.170	−33.6
Helium	0.179	−268.7
Hydrogen	0.09	−252.7
Hydrogen Chloride	1.630	−82.9
Hydrogen Fluoride	0.893	19.4
Hydrogen Sulfide	1.518	−61.6
Methane	0.714	−164.0
Nitric Oxide	1.339	−153.0
Nitrogen	1.250	−195.7
Nitrous Oxide	1.964	−89.8
Oxygen	1.43	−182.9
Sulfur Dioxide	2.857	−10.1

DISPLACEMENT SERIES

1. Potassium
2. Sodium
3. Magnesium
4. Aluminum
5. Manganese
6. Zinc
7. Chromium
8. Cadmium
9. Iron
10. Cobalt
11. Nickel
12. Tin
13. Lead
14. *Hydrogen*
15. Copper
16. Arsenic
17. Bismuth
18. Antimony
19. Mercury
20. Silver
21. Palladium
22. Platinum
23. Gold

SOLUBILITY OF VARIOUS GASES IN WATER

Name of Gas	Volume Absorbed, at 0° C. and 760 mm. Pressure, by 1 Liter of Water
Ammonia	1298.9000 liters
Hydrogen Chloride	506.0000 liters
Sulphur Dioxide	79.7900 liters
Hydrogen Sulphide	4.3700 liters
Carbon Dioxide	1.7130 liters
Oxygen	0.0496 liter
Nitrogen	0.0233 liter
Hydrogen	0.0214 liter

SOLUBILITY OF VARIOUS SOLIDS IN WATER

Substance	Formula	Weight Dissolved by 100 cc. of Water at		
		0° C.	20° C.	100° C.
Calcium Chloride	$CaCl_2$	49.500 g.	74.500 g.	159.000 g.
Sodium Chloride	NaCl	35.700 g.	36.000 g.	39.800 g.
Potassium Nitrate	KNO_3	13.300 g.	31.600 g.	246.000 g.
Copper Sulphate	$CuSO_4$	14.300 g.	21.700 g.	75.400 g.
Calcium Sulphate	$CaSO_4$	0.759 g.	0.203 g.	0.162 g.
Calcium Hydroxide	$Ca(OH)_2$	0.185 g.	0.156 g.	0.077 g.

THE METRIC SYSTEM

The metric system is a decimal system throughout. The different units, with their most important subdivisions and multiples, are as follows:

Linear Measure

Metric Units			Equivalents in English Units
1 millimeter (mm.)	0.001	meter (m.)	0.03937 in.
1 centimeter (cm.)	0.01	meter	0.3937 in.
1 decimeter (dm.)	0.1	meter	3.937 in.
	1.0	meter	39.37 in.
1 dekameter	10.0	meters	
1 hektometer	100.0	meters	
1 kilometer	1000.0	meters	0.6214 mi.

Volume

Metric Units			Equivalents in English Units
1 cubic centimeter	(cc.)	1 centimeter cubed	0.0338 fl. oz. (U. S.) 0.0352 fl. oz. (British)
1 liter	(l.)	1000 cubic centimeters	1.057 quarts (U. S.) 1.136 quarts (British)

Weight

The unit of weight is the gram, which may be defined as the mass of one cubic centimeter of water at 4° C. The value is determined at this temperature because at 4° C. water has its greatest density. The subdivisions and multiples of the gram are as follows:

Metric				English
1 milligram	(mg.)	0.001	gram (g.)	
1 centigram	(cg.)	0.01	gram	
1 decigram	(dg.)	0.1	gram	
		1.0	gram	0.0353 oz. Av.
1 dekagram		10.0	grams	
1 hektogram		100.0	grams	
1 kilogram		1000.0	grams	2.204 lb. Av.

PRESSURE EXERTED BY AQUEOUS VAPOR, 0° TO 50° C.

Expressed in Millimeters of Mercury

°C.	0.0	0.2	0.4	0.6	0.8	°C.	0.0	0.2	0.4	0.6	0.8
0	4.57	4.64	4.70	4.77	4.84	26	24.96	25.25	25.55	25.85	26.16
1	4.91	4.98	5.05	5.12	5.20	27	26.47	26.78	27.10	27.42	27.74
2	5.27	5.35	5.42	5.50	5.58	28	28.07	28.39	28.73	29.06	29.40
3	5.66	5.74	5.82	5.90	5.99	29	29.74	30.09	30.44	30.79	31.15
4	6.07	6.15	6.24	6.33	6.42	30	31.51	31.87	32.34	32,61	32.99
5	6.51	6.60	6.69	6.78	6.88	31	33.37	33.75	34.14	34.53	34.92
6	6.97	7.07	7.17	7.26	7.36	32	35.32	35.72	36.13	36.54	35.95
7	7.47	7.57	7.67	7.78	7.88	33	37.37	37.79	38.22	38.65	39.08
8	7.99	8.10	8.21	8.32	8.43	34	39.52	39.97	40.41	40.87	41.32
9	8.55	8.66	8.78	8.90	9.02	35	41.78	42.25	42.72	43.20	43.67
10	9.14	9.26	9.39	9.51	9.64	36	44.16	44.65	45.14	45.64	46.14
11	9.77	9.90	10.03	10.16	10.30	37	46.65	47.16	47.68	48.20	48.73
12	10.43	10.57	10.71	10.85	10.99	38	49.26	49.80	50.34	50.89	51.44
13	11.14	11.28	11.43	11.58	11.73	39	52.00	52.56	53.13	53.70	54.28
14	11.88	12.04	12.19	12.35	12.51	40	54.87	55.46	56.05	56.65	57.26
15	12.67	12.84	13.00	13.17	13.34	41	57.87	58.49	59.11	59.74	60.34
16	13.51	13.68	13.86	14.04	14.21	42	61.02	61.66	62.32	62.98	63.64
17	14.40	14.58	14.76	14.95	15.14	43	64.31	64.99	65.67	66.36	67.06
18	15.33	15.52	15.72	15.92	16.12	44	67.76	68.47	69.18	69.90	70.63
19	16.32	16.52	16.73	16.94	17.15	45	71.36	72.10	72.85	73.60	74.36
20	17.36	17.58	18.00	18.02	18.24	46	75.13	75.91	76.69	77.48	78.27
21	18.47	18.69	18.92	19.16	19.40	47	79.07	79.88	80.70	81.52	82.35
22	19.63	19.87	20.11	20.36	20.61	48	83.19	84.03	84.89	85.75	86.61
23	20.86	21.11	21.37	21.63	21.89	49	87.49	88.37	89.26	90.16	91.06
24	22.15	22.42	22.69	22.96	23.24	50	91.98	92.90	93.83	94.77	95.71
25	23.52	23.80	24.08	24.37	24.66						

A PROBLEM COURSE IN CHEMISTRY

CHAPTER I. PRELIMINARY CALCULATIONS

Metric Measurements

The metric system of weights and measures is universally used in scientific work. The fundamental unit of this system is the *meter*, which may be defined as the distance between two parallel scratches placed near the ends of a bar of metal (platinum-iridium alloy) which is kept in the vaults of the International Bureau of Weights and Measures at Sèvres, a suburb of Paris, France. The measurement must be taken when the temperature is at 0° C.

This distance was originally intended to represent one ten-millionth part of the distance from the equator to the north pole, measured along the meridian passing through Paris, but a slight error was made in the calculations. However, for many good reasons, the standard has never been adjusted to the value later determined to be accurate for the distance from the equator to the north pole.

Similar standard meters are kept in London, England, and at the United States Bureau of Standards in Washington, D. C. Duplicates are located at many seats of government throughout the world.

1. A rod measures 15 centimeters in length. One third of it is cut off. What is the length in centimeters of the piece remaining? In millimeters?

2. Some pieces of glass tubing were measured and their respective lengths were found to be: 45.8 cm., 84.6 cm., 29.7 cm., 6.1 cm., and 63.8 cm. Find the combined length in meters.

3. The top of a table is 90 cm. long and 60 cm. wide. What would be the area in square centimeters? Find the area in square meters.

4. In measuring certain stretches of track the following lengths were recorded: 735.2 m., 289.7 m., 900 m., 582.4 m., 899.5 m., and 2374.7 m. Find the combined length in kilometers.

5. A vessel measures 45 cm. by 60 cm. by 20 cm. How many cubic centimeters of water would it hold?

6. 150 cubic centimeters of water are poured into a vessel, the capacity of which is exactly 1 liter. What volume, expressed in cubic centimeters, remains unoccupied?

7. A room is 15 m. long, 9 m. wide, and 4 m. high. Find the total surface in the walls and ceiling, counting doors and windows. What would be the volume of the air in the room in cubic meters?

8. 1 liter of air weighs 1.293 grams. Find the total weight in kilograms of the air in the room mentioned in the preceding problem.

9. A stone measures 2.5 m. in length, 80 cm. in width, and 25 cm. in thickness. Its specific gravity is 2.6. Compute the total weight of the stone and the pressure in grams per square centimeter, while lying on its edge.

10. 1 liter of air weighs 1.293 grams. What is the weight of 1 cubic meter of air expressed in kilograms?

11. 1 cubic centimeter of water at 4° C. has a mass of 1 gram. What is the weight in grams of 1.28 liters of water at this temperature?

12. A swimming tank measures 17 m. in length, 9 m. in width, and 1.75 m. in average depth. It is filled with sea water which has a density of 1.027. Find the total volume of the tank in liters, and the total weight of the water in kilograms.

Specific Gravity

By the *density* of a substance is meant the number of units of mass in a given unit of volume of the substance. *Example:* The density of aluminum is 2.7 g. per cubic centimeter. The density of water (at 4° C.) is 1 g. per cubic centimeter, or 62.4 lb. per cubic foot. The density of oxygen is 1.429 g. per liter.

When the density of a substance is found relative to that of another substance which is taken as a standard, the term *specific gravity* is applied to the number which expresses the ratio between them.

The specific gravity of solids and of liquids is determined in comparison with water as a standard, while that of a gas is usually referred to air as a standard.

As an example, 1 cc. of iron weighs 7.56 g. 1 cc. of water weighs 1 g. The ratio is 7.56 : 1, or $\frac{7.56}{1} = 7.56$. This number represents the specific gravity of the iron.

It will be noticed that, since 1 cc. of water weighs 1 g. and is always the denominator in the fraction indicating the specific gravity of a solid or a liquid, the weight of 1 cc. of any solid or liquid substance expressed in grams will indicate directly the specific gravity of the substance. Such relationships will not hold good in the English system of weights and measures.

Since gases are so much lighter than water, they are usually referred to a gas as a standard. Air is commonly the standard. The specific gravity of a gas is then the density of the gas compared with that of air measured under the same conditions.

For example, 1 liter of air under standard conditions of temperature and pressure weighs 1.293 grams. 1 liter of oxygen under the same conditions weighs 1.429 grams. Therefore, the specific gravity of oxygen is $\frac{1.429}{1.293} = 1.1052$.

1. 250 cubic centimeters of a liquid weigh 300 grams. What is the relative density of the liquid, compared with that of water?

2. Pure iron has a specific gravity of 7.28. What will be the weight in grams of a block of iron with a volume of 132 cubic centimeters?

3. A cube of marble which measures 1.5 centimeters on each side weighs 9.45 grams. What is its specific gravity?

4. What volume of copper, in cubic centimeters, is required to balance 100 grams of lead, if the specific gravity of the copper is 8.95 and the specific gravity of the lead is 11.3?

Fundamentals of Chemical Calculations

When any chemical change takes place, the amounts of materials involved bear definite ratios to each other, by weight. For any given chemical changes, these ratios are invariable. One gram of carbon in burning to form carbon dioxide always combines chemically with $2\frac{2}{3}$ grams of oxygen to form $3\frac{2}{3}$ grams of carbon dioxide.

In order to determine what are the ratios in which substances react chemically with each other, we make use of chemical equations, which tell us definitely what substances are in reaction with each other, what substances are formed, and, relatively, what amounts of these substances are actively present.

Since chemical elements are simple substances which cannot be decomposed by any known means, unit quantities of these elements serve as the basis of our equations. These unit quantities of the elements are the smallest particles of the elements which can enter into chemical combination, and are called *atoms*.

All the atoms of a given kind have the same characteristic weight. Atoms of different kinds always have different weights. These weights are not determined absolutely, since the atoms are too small to be observed individually. However, the relative weights of different kinds of atoms may be determined accurately. In order to serve as a basis of comparison of these relative weights, one weight must be taken as a standard. For many reasons the weight of the oxygen atom is chosen as the most convenient standard, and to it is assigned the value 16.

The atomic unit of weight might then be said to be $\frac{1}{16}$ of the weight of an atom of oxygen. If one atom of an element, sulfur, for example, weighs 32 of these units, that is, $\frac{32}{16}$ of the weight of one atom of oxygen, it is given the value 32 for its atomic weight. One atom of mercury is found to weigh 200 of these units, which means that it has $\frac{200}{16}$ of the weight of one atom of oxygen. Therefore we say that mercury has an atomic weight of 200.

A table of the weights of the atoms is given on page xi.

To distinguish the atoms in our equations, we make use of *symbols* which represent the alphabet of a system of chemical shorthand. Thus, **C** represents carbon, one atom; **O** represents oxygen, one atom; **Fe** represents iron, one atom; and **S** represents sulfur, one atom.

When iron (symbol **Fe**) and sulfur (symbol **S**) are heated together, they unite chemically to form a new substance which has properties distinctly its own, resembling neither the iron nor the sulfur from which it is made. This new substance, since it contains elements in chemical combination with each other, is called a *chemical compound*; this particular compound is called *iron sulfide*. It has been found by careful experiment that just one atom of iron unites chemically with one atom of sulfur to form iron sulfide. The compound formed may be represented thus: **FeS**. Such a group of symbols written together is called a *formula*. This formula indicates the least amounts of the elements that can enter into chemical combination to form iron sulfide; consequently the formula represents the least amount of iron sulfide that can exist. This amount of the compound is called a *molecule*.

The equation for the change is now written:

$$\mathrm{Fe} + \mathrm{S} \longrightarrow \mathrm{FeS}.$$

This is merely a brief way of saying that, when heated together, one atom of iron unites chemically with one atom of sulfur to form one molecule of iron sulfide.

By reference to our table of atomic weights, we find that iron has an atomic weight of 56, while sulfur has an atomic weight of 32. Our equation tells us, then, that 56 parts by weight of iron combine chemically with 32 parts by weight of sulfur, to form 88 parts by weight of iron sulfide. Iron and sulfur always combine in these proportions to form this compound.

If we have, say, 10 grams of iron and wish to know how much sulfur is required to combine with it to form iron sulfide, we may write our equation thus, indicating the atomic weights over the atoms and the gram weights under the substances:

$$\underset{10\text{ g.}}{\overset{56}{\mathrm{Fe}}} + \underset{x\text{ g.}}{\overset{32}{\mathrm{S}}} \longrightarrow \overset{88}{\mathrm{FeS}}.$$

56 parts of iron combine with 32 parts of sulfur by weight. Therefore, 10 grams of iron combine with the same proportionate number of grams of sulfur:

$$56 : 32 :: 10\text{ g.} : x,$$

x being the number of grams of sulfur required. Since in any true proportion the product of the means is equal to the product of the extremes, we have:

$$56\,x = 320\text{ grams}$$

or

$$x = \frac{320\text{ grams}}{56}$$

$$= 5.7\text{ grams of sulfur required.}$$

Relation of Volume to Weight in Gas Problems

It is often necessary to calculate the volume of one or more gaseous products obtained from a given weight of a substance, or to find the weight of one or more substances required for the production of a given amount by volume of a gaseous substance. To illustrate this point we might take as an example a typical problem in the following: How many liters of oxygen could be obtained by decomposing 30 grams of potassium chlorate, $KClO_3$? The equation for the reaction is:

$$2\,\underset{30}{\overset{245}{\mathrm{KClO_3}}} \longrightarrow 2\,\mathrm{KCl} + 3\,\underset{x}{\overset{96}{\mathrm{O_2}}}.$$

We find that the combined weights of the atoms in 2 molecules of potassium chlorate, $KClO_3$, equal ?

Acetylene
Air
Ammonia
Argon
Carbon Dioxide
Carbon Monoxide
Chlorine
Helium
Hydrogen
Hydrogen Chloride
Hydrogen Bromide
Hydrogen iodide
Methane
Nitric Oxide
Nitrogen
Nitrous Oxide
Oxygen
Sulfur Dioxide

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eated until it
rcury, Hg, was
O_2, present in the

O_2.

4. A small amount of manganese dioxide, MnO_2, was mixed with 122.5 grams of potassium chlorate, $KClO_3$, and the mixture was heated in a test tube until there was no further change. The gas evolved was collected over water. What gas was evolved and how many liters could be collected? The equation is:

$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2.$$

1 liter of oxygen weighs 1.43 grams.

5. How many grams of potassium chlorate, $KClO_3$, must be decomposed by heating with a little manganese dioxide, MnO_2, to produce 25 liters of gas? The reaction is:

$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2.$$

6. 39.2% of potassium chlorate, $KClO_3$, is oxygen. How much potassium chlorate by weight would be needed to produce 50 liters of oxygen, O_2, if 1 liter of oxygen weighs 1.43 grams? The equation for the reaction is:

$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2.$$

7. In the electrolysis of 15.1 grams of water, H_2O, how many 500 cc. bottles could be filled with oxygen, O_2? Equation:

$$2\ H_2O \longrightarrow 2\ H_2 + O_2.$$

8. How many bottles, each containing 300 cc., could be filled with the oxygen, O_2, obtained from 30 g. of potassium chlorate, $KClO_3$, allowing for 20% wastage in collecting the gas? The reaction is expressed by:

$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2.$$

9. 1 liter of oxygen, O_2, weighs 1.43 grams. What is the relative density of oxygen compared with air?

10. What weight of mercury, Hg, will be produced when 10 grams of mercuric oxide, HgO, is decomposed in the preparation of oxygen? The equation for the reaction is:

$$2\ HgO \longrightarrow 2\ Hg + O_2.$$

Percentage Composition of Compounds

We have already learned that the weight of a molecule of a compound is the sum of the weights of the atoms in it. If we wish to determine what per cent part of the molecule is due to each of the elements composing it, we may divide the total atomic weight present of each element by the weight of the molecule. The result in each case is the per cent part of the element in the compound.

Example: Calculate the percentage composition of crystallized copper sulfate, $CuSO_4 \cdot 5\,H_2O$.

Solution:

$$\frac{249.6\,.}{\underset{Cu}{63.6}\ \ \underset{S}{32}\ \ \underset{O_4\cdot}{64}\ \ \underset{5H_2O}{10\ \ 80}}.$$

The atomic weight of copper present is seen to be 63.6; of sulfur, 32; of oxygen, 64+80, or 144; and of hydrogen, 10. The sum of these weights, which is the molecular weight of crystallized copper sulfate, is 249.6.

We then have:

$$\begin{aligned} 63.6/249.6 &= 25.5\%\ \text{copper} \\ 32/249.6 &= 12.8\%\ \text{sulfur} \\ 144/249.6 &= 57.7\%\ \text{oxygen} \\ 10/249.6 &= 4.0\%\ \text{hydrogen} \end{aligned}$$

1. What is the percentage composition of potassium chlorate which has the formula, $KClO_3$?

2. What weight of mercury, Hg, will be produced when 10 grams of oxygen, O_2, are evolved by decomposing mercuric oxide, HgO, according to the equation:

$$2\,HgO \longrightarrow 2\,Hg + O_2?$$

3. What weight of potassium chloride, KCl, would remain after 20 grams of potassium chlorate, $KClO_3$, had been heated until no further oxygen, O_2, was given off? Equation:

$$2\,KClO_3 \longrightarrow 2\,KCl + 3\,O_2.$$

CHAPTER III. HYDROGEN

Hydrogen

Hydrogen, the lightest of the elements, occurs in water, in all acids, and in all living tissue. It may be prepared in the pure form by the electrolysis of water or by its displacement from an acid by a metal such as zinc. Hydrogen makes up a large part of most fuel gases.

1. 2 grams of pure hydrogen, H_2, was obtained by the action of sodium, Na, upon water, H_2O. Calculate the number of grams of sodium required to produce it. The equation is:

$$2\,Na + 2\,H_2O \longrightarrow 2\,NaOH + H_2.$$

2. How many grams of hydrogen, H_2, could be obtained by the action of zinc, Zn, upon 98 grams of sulfuric acid, H_2SO_4? What volume would this gas occupy in liters? This reaction is expressed by the equation:

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2.$$

3. Water, H_2O, containing a small amount of sulfuric acid, H_2SO_4, was decomposed by passing a current of electricity through it and the evolved gases were collected and weighed. If the oxygen gas, O_2, weighed 320 grams, what would be the weight of the hydrogen, H_2? The equation for the reaction is:

$$2\,H_2O \longrightarrow 2\,H_2 + O_2.$$

4. 1 gram of pure sodium, Na, is placed in water, H_2O, and all the hydrogen, H_2, produced during the reaction is collected and weighed. What is the weight of the hydrogen produced? The reaction is:

$$2\ Na + 2\ H_2O \longrightarrow 2\ NaOH + H_2.$$

How much water is decomposed?

5. How many grams of dilute sulfuric acid, laboratory strength 20% H_2SO_4, must be used to produce 3 grams of hydrogen, H_2, by its reaction with zinc, Zn? How much zinc is used up? The reaction is as follows:

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2.$$

6. Water, H_2O, is decomposed by passing an electric current through a weak acid or basic solution. The hydrogen, H_2, produced is found to occupy a volume of 32.6 cc. What is the corresponding volume of oxygen, O_2, produced? The reaction is expressed by:

$$2\ H_2O \longrightarrow 2\ H_2 + O_2.$$

7. 20 grams of a mixture of zinc, Zn, and sand was mixed with dilute hydrochloric acid, HCl, and the action was allowed to continue until there was no zinc left. After separating, drying, and weighing, the zinc chloride, $ZnCl_2$, formed was found to have a weight of 25 grams. What per cent by weight of the original mixture was sand?

$$Zn + 2\ HCl \longrightarrow H_2 + ZnCl_2.$$

8. How many grams of water would be used up in the reaction with 30 grams of sodium, Na, to produce sodium hydroxide, NaOH, and hydrogen? Equation is:

$$2\ Na + 2\ H_2O \longrightarrow 2\ NaOH + H_2.$$

9. In the electrolysis of 16.1 grams of water it was found that 10 liters of oxygen gas was given off. From this calculate the percentage of hydrogen in water. Equation:

$$2\ H_2O \longrightarrow 2\ H_2 + O_2.$$

Meaning of Per Cent Strength

It is very frequently necessary in connection with these problems to state the per cent strength of a substance. Wherever this is done, the percentage is based upon weight composition unless otherwise specifically stated. For example, we may state the per cent strength of a solution of sodium hydroxide as 30% NaOH. This would mean that in each 100 grams of sodium hydroxide solution taken there would be 30 grams of pure NaOH. When we speak of sulfuric acid, 98% H_2SO_4, we mean that in each 100 grams of sulfuric acid solution there are 98 grams of pure H_2SO_4.

1. What weight of hydrogen could be obtained from 100 grams of dilute hydrochloric acid, assuming that 10% of the hydrochloric acid is HCl?

2. How much sulfuric acid, 20% H_2SO_4, by weight would be required to produce 250 liters of hydrogen?

3. Taking the cost of mossy or granulated zinc, Zn, as 90 cents per pound, what would be the cost of the zinc required to produce 500 liters of hydrogen, by its reaction with an acid?

4. How much would be saved in the process of the preceding problem if iron filings, Fe, priced at 10 cents per pound, were used in place of the zinc?

5. A dirigible balloon that had a capacity of 7000 cubic meters was filled with hydrogen. What would be the cost of the sulfuric acid that would be required to furnish the hydrogen to fill it? Assume that commercial sulfuric acid, 96% H_2SO_4, was used and that it cost 2 cents per pound. 1 pound is equivalent to 453.6 grams. The reaction is:

$$Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2.$$

6. How many liters of hydrogen would be formed by passing 50 grams of water in the form of steam over highly heated iron filings, Fe? The reaction is expressed:

$$3\ Fe + 4\ H_2O \longrightarrow FeO \cdot Fe_2O_3 + 4\ H_2.$$

7. The specific gravity of hydrogen gas, compared with air, is 0.0697. What is the weight of 1 liter of hydrogen, expressed in grams, to the second decimal?

8. What volume, in liters, of hydrogen could be produced from 150 grams of dilute sulfuric acid, 20% H_2SO_4?

9. What is the cost per liter of hydrogen gas if zinc sells at $2 per kilogram and sulfuric acid, 95 % H_2SO_4, sells at $0.20 per kilogram?

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2.$$

10. How many pounds of iron would be required to displace all of the hydrogen of 100 pounds of hydrochloric acid which is 35% HCl?

$$Fe + 2\ HCl \longrightarrow FeCl_2 + H_2.$$

11. What would be the cost of the magnesium, Mg, necessary to release 1 liter of hydrogen from hydrochloric acid, HCl, if the magnesium costs $0.09 per gram?

$$Mg + 2\ HCl \longrightarrow MgCl_2 + H_2.$$

12. How many grams of fine iron wire, Fe, heated to a high temperature, will be required to produce 10 liters of hydrogen by its action upon steam?

$$3\ Fe + 4\ H_2O \longrightarrow FeO \cdot Fe_2O_3 + 4\ H_2.$$

13. What is the relative density of oxygen which weighs 1.43 grams per liter, compared with hydrogen which weighs 0.09 gram per liter?

CHAPTER IV. HYDROGEN-OXYGEN COMPOUNDS

Water

Water, the universal solvent, is very difficult to produce in a perfect degree of purity. The various means of purifying natural water depend upon the use to which it is to be put, the method used in each case being the cheapest that will answer the purpose.

Determinations upon the composition of water show it to contain 88.81 % oxygen. From this one would naturally suppose water to be a good oxidizing agent, but such is not the case. Water is such a stable compound that it will surrender its oxygen only with difficulty.

1. A mixture of oxygen and hydrogen in a eudiometer tube was exploded by an electric spark, forming water. The following data on the experiment were taken:

Volume of oxygen	17.5 cc.
Volume of hydrogen and oxygen	60. cc.
Volume of gas remaining after explosion . . .	7.5 cc.

The residual gas was found to be combustible, but would not support combustion. From these results calculate the ratio by volume in which hydrogen and oxygen combine to form water.

2. Pure hydrogen gas was passed through a heated tube containing dried copper oxide, CuO, and the water that was formed was collected in a U-tube. If the increase in weight of the U-tube was 20 grams, what was the loss in weight of the tube containing the copper oxide? The equation for the reaction is:

$$H_2 + CuO \longrightarrow H_2O + Cu.$$

3. A mixture of air and hydrogen in a eudiometer tube was exploded and the contraction in volume noted. The following data were obtained:

Volume of air	50 cc.
Volume of hydrogen and air	80 cc.
Volume of gas remaining after explosion . . .	50 cc.

Calculate the percentage of oxygen in the sample of air tested.

4. Ten liters of hydrogen was passed through a tube containing highly heated copper oxide, CuO. What weight of water was formed in the reaction? The equation for the reaction is as follows:

$$H_2 + CuO \longrightarrow H_2O + Cu.$$

Definite Composition

Careful analysis of a pure compound, no matter how often it is made, invariably gives the same percentage composition, within the limits of unavoidable error. This fact may be briefly stated as the Law of Definite Composition: "The composition of a chemical compound never varies."

1. Upon analyzing some pure copper chloride, $CuCl_2$, 30 grams of copper, Cu, was found present. What was the weight of the original copper chloride?

2. In an analysis of 14 kilograms of lead oxide, PbO, what weight of lead, Pb, would be found present?

Hydrogen Peroxide

Hydrogen peroxide is an unstable compound which breaks down easily, with the formation of water and free oxygen. Consequently, it is a good oxidizing agent in solution.

1. How many pounds of hydrogen peroxide, 3% H_2O_2, can be made from 153 pounds of barium peroxide, BaO_2? The equation is:

$$BaO_2 + H_2SO_4 \longrightarrow BaSO_4 + H_2O_2.$$

2. How many grams of barium sulfate, $BaSO_4$, will be formed in the preparation of 100 grams of hydrogen peroxide, 3% H_2O_2, from barium peroxide, BaO_2, and sulfuric acid, H_2SO_4? The equation for the reaction is:

$$BaO_2 + H_2SO_4 \longrightarrow BaSO_4 + H_2O_2.$$

3. What weight of hydrogen peroxide, 5% H_2O_2, could be prepared from 2 kilograms of barium peroxide, BaO_2, by the action of sulfuric acid, H_2SO_4, upon it? The reaction is as follows:

$$BaO_2 + H_2SO_4 \longrightarrow BaSO_4 + H_2O_2.$$

4. Calculate the volume of oxygen that could be obtained from 8 kilos of hydrogen peroxide, 3% H_2O_2, for sanative purposes.

$$2\ H_2O_2 \longrightarrow 2\ H_2O + O_2.$$

5. How many grams of sulfuric acid solution, 25% H_2SO_4, will be required to combine completely with 350 grams of barium peroxide, BaO_2, according to the reaction expressed as follows:

$$BaO_2 + H_2SO_4 \longrightarrow BaSO_4 + H_2O_2.$$

6. How many kilograms of hydrogen peroxide, H_2O_2, will be required to produce 12 liters of oxygen when completely decomposed by boiling? The equation is:

$$2\ H_2O_2 \longrightarrow 2\ H_2O + O_2.$$

CHAPTER V. HEAT MEASUREMENT

In the measurement of the degree of heat, the Centigrade scale of the thermometer is used. In this scale, the standard points are the freezing point of water, which is marked 0° C. on the scale, and the boiling point of water under a pressure of 760 mm. of mercury, which is marked 100° C. The range of temperature between these points is divided into 100 equal divisions or degrees.

The thermometer ordinarily used in the household in the United States and in Great Britain is graduated according to the Fahrenheit scale, in which 32° represents the freezing point of water, and 212° the boiling point of water. The range between these points is divided into 180 equal divisions or degrees.

Since 100 degrees C. covers the same range as 180 degrees F., each degree C. is equal to $\frac{180}{100}$ of one degree F., or $\frac{9}{5}$ of one degree F.

Also, each degree F. is equal to $\frac{100}{180}$ of one degree C., or $\frac{5}{9}$ degree C.

It will be noticed that 32° F. represents the same temperature as 0° C. Accordingly, to change a reading in Fahrenheit degrees to a reading in Centigrade degrees, one first subtracts 32 from the Fahrenheit reading, then multiplies the remainder by $\frac{5}{9}$.

To change from Centigrade reading to Fahrenheit reading, one first multiplies the Centigrade reading by $\frac{9}{5}$, then adds 32 degrees.

The thermometer serves to indicate the degree of intensity of heat. When it comes to measurement of amount of heat, further units than the degree are necessary, since the amount of heat present in any case depends upon the amount of material considered and upon the kind of material. Water has the greatest heat capacity of all substances; that is, it requires more heat to raise the temperature of a given amount of water through a given range of temperature than it does for any other substance.

The amount of heat required to raise the temperature of 1 gram of water through 1 degree C. is taken as the standard unit for heat measurement, and is called the *calorie*. To raise the temperature of 10 grams of water through 1 degree C. requires, of course, 10 calories of heat. To raise the temperature of 1 gram of water through 10 degrees C. requires 10 calories of heat. In other words, to determine the amount of heat required to raise the temperature of any number of grams of water through any number of degrees C., we multiply the number of grams by the number of degrees C. rise in temperature. The result is calories required.

Example: How many calories of heat are required to raise the temperature of 16 grams of water from 18° C. to 36° C.?

Solution: $36° - 18°$ is 18 degrees rise in temperature. $18 \times 16 =$ *288 calories* required.

Water in cooling gives out the same number of calories of heat as the same amount of water absorbs in having its temperature raised through an equal range in temperature. Therefore, to find the quantity of heat given out by an amount of water in cooling, we multiply the weight of the water in grams by the fall in temperature in degrees C. The result is calories of heat given out.

Example: How many calories of heat are given out by 150 grams of water in cooling from 100° C. to 40° C.?

Solution: 100−40 is equal to 60. 60×150 is equal to *9000 calories*, the amount of heat given out.

When a substance in the solid state changes to the liquid state, a new type of molecular motion occurs in the substance. Such a change in type of molecular motion results only from a change in the amount of heat energy present in the substance. This heat energy is not used to change the temperature of the material, merely to make the molecules move at a different average velocity and in a new type of arrangement.

When a change in state occurs, we find that a solid in going to a liquid, or a liquid in going to a gas or vapor, absorbs heat. On the other hand, when a gas or vapor condenses to a liquid this heat energy is given off, as well as when a liquid freezes to a solid.

The amount of heat in calories required to convert one gram of a solid to one gram of a liquid is called the *heat of fusion*. Incidentally, when the liquid solidifies, the same amount of heat is given out by each gram of material, and is then called *heat of solidification*. In order to change one gram of ice at 0° C. to water at 0° C., 80 calories of heat must be supplied to the ice. *80 calories*, then, is the heat of fusion of ice.

When a liquid in turn is converted into a vapor or a gas, still another type of molecular motion occurs within the substance. The molecules move with much greater velocity, the average distance between the molecules becomes much greater, and the molecules move freely in all directions unless turned back by obstructions. This change requires a large amount of energy, which in this case is called the *heat of vaporization*. It requires 537 calories of heat to convert one gram of water at 100° C. to steam at 100° C. *537 calories* is, then, the heat of vaporization of water.

1. How many calories of heat are required to raise the temperature of 25 grams of water from 0° C. to 100° C.?

2. In a certain experiment the temperature of 3000 grams of water is raised from 21.9° C. to 26.5° C. by the combustion of 3 grams of material. How much heat is produced by the combustion of 1 gram of this material?

3. 50 grams of water at 20° C. is mixed with 150 grams of water at 80° C. Disregarding heat absorbed by the vessels containing the water, and heat lost by radiation, what is the final temperature of the mixture?

4. 1 gram of hydrogen when burned heats 1 liter of water from 20° C. to 48.8° C. How much heat is produced by burning this amount of hydrogen?

5. How many calories of heat would be needed to convert 4000 grams of ice at – 40° C. to steam at 150° C.? Assume that the specific heat of ice and of steam is 0.5, which means that only 0.5 calories of heat would be needed to produce a change of 1 degree C. in the temperature of 1 gram of the steam or of the ice.

6. How many grams of pure carbon would be required to raise the temperature of 40 liters of water from 10° C. to 95° C., through its heat of combustion? 1 gram of carbon in burning gives out 8040 calories of heat.

7. How many calories of heat are required to raise the temperature of 156 grams of water from 16° C. to 38° C.?

8. 50 grams of water at 90° C. is mixed with 100 grams of water at 15° C. What is the final temperature of the mixture?

9. 1 gram of hydrogen in burning completely releases 28,800 calories of heat. If 50% of the heat released by burning 10 liters of hydrogen is applied to raising the temperature of 1 kilogram of water, original temperature 20° C., what will be the final temperature of the water?

CHAPTER VI. WRITING AND BALANCING OF EQUATIONS

The equation is the shorthand of chemistry. In describing a certain reaction we might take several whole paragraphs to explain more or less fully the nature and amounts of the reacting materials, the conditions under which they react, the nature and amounts of products. This would be cumbersome and ineffective. For brevity, as well as for accuracy, the equation has been developed as a means of expression.

We should, however, remember that the equation describes a reaction that has been found by experiment actually to have taken place and that, since this is true, one must first understand the reaction involved before he can intelligently undertake the writing of the equation describing it. To write an equation describing a chemical reaction one must first know the following:

(1) The substances that take part in the reaction;
(2) The composition of each reacting substance;
(3) All the products that are formed;
(4) The composition of each product.

With these facts in hand one can proceed to write the equation for the reaction.

For instance, we may wish to describe the reaction taking place when steam is passed over highly heated iron filings. By experiment we find that, if steam is passed over highly heated iron filings, hydrogen is given off while iron oxide is left in the tube. Knowing these facts and also the composition of all the materials involved, we can write an equation for the reaction as follows:

$$H_2O + Fe \longrightarrow H_2 + Fe_3O_4. \quad \text{(Unbalanced)}$$

This type of equation serves as a qualitative description of the reaction, but does not give any of the amounts of the materials involved. The fundamental principle upon which chemical calculations depend is the indestructibility of matter, so the equation must represent the same amount of each element after the change as before. There must be the same number of atoms of each element represented on each side of the equation.

In the equation as it stands we find that there are three atoms of *iron* in the iron oxide, Fe_3O_4, while there is only one atom of iron, Fe, on the left side of the equation. The obvious thing to do then would be to multiply the one molecule containing one atom of iron by three in order to obtain three molecules each containing one atom of iron. This one change would make our equation read:

$$H_2O + 3\ Fe \longrightarrow H_2 + Fe_3O_4. \quad \text{(Unbalanced)}$$

We still find that the number of *oxygen* atoms is not the same on the two sides of the equation. There are four atoms of oxygen on the right side of the equation in the molecule of iron oxide, while there is only one atom of oxygen on the left side of the equation in the one molecule of steam. We can then multiply the one molecule of steam by four in order to get four molecules of steam each supplying one atom of oxygen. This change makes our equation read:

$$4\ H_2O + 3\ Fe \longrightarrow H_2 + Fe_3O_4. \quad \text{(Unbalanced)}$$

The only element that is not balanced now is the *hydrogen.*

There is one molecule containing two atoms of hydrogen on the right side of the equation, while there are four molecules of steam each containing two atoms of hydrogen on the left side of the equation. We only need then to multiply the one molecule of hydrogen on the right side by four to obtain the complete balanced equation as follows:

$$4\,H_2O + 3\,Fe \longrightarrow 4\,H_2 + Fe_3O_4. \quad \text{(Balanced)}$$

This equation now states three distinct facts in a concise way.

1. Qualitatively, it states that hydrogen and iron oxide are formed in the reaction between water and iron.

2. Quantitatively, it states that 4 formula weights of water (4×18.016) react with 3 formula weights of iron (3×55.9) to form 4 formula weights of hydrogen (4×2.016) and 1 formula weight of iron oxide, 231.7.

3. Molecularly, it states that 4 molecules of water react with 3 molecules of iron to form 4 molecules of hydrogen and 1 molecule of iron oxide.

We should always remember that the equation is an expression of a reaction which we know to have taken place and is not, as a rule, written before the reaction itself is studied.

1. Complete the following equations:

$KClO_3$ (when heated) $\longrightarrow$ . . . .
$Zn + H_2SO_4 \longrightarrow$ . . . .
$Mg +$ . . . $\longrightarrow MgCl_2 +$ . . . . . .
$NaOH + HCl \longrightarrow H_2O +$ . . . .

2. Complete the following equations:

$Ca(OH)_2 +$ . . . $\longrightarrow Ca(NO_3)_2 +$ . .
$KOH + H_2SO_4 \longrightarrow H_2O +$ . . . .
$H_2CO_3 + NaOH \longrightarrow$ . . . $+ Na_2CO_3$.
$H_2SO_4 +$ . . . $\longrightarrow H_2O + Na_2SO_4$.

CHAPTER VII. NITROGEN AND THE ATMOSPHERE

Nitrogen

Nitrogen, a generally sluggish element, is difficult to force into combination and leaves its compounds with great ease. Its compounds are, therefore, unstable. Nitrogen is found in all protein tissue and is widely used in the making of explosives. The high cost of nitrogen fertilizers is due to the scarcity of nitrogen, in combined form. Nitrogen compounds have been built up with extreme slowness in nature and are at present being consumed faster than nature can produce them.

1. How many grams of ammonium nitrite, NH_4NO_2, would be required to produce 28 liters of nitrogen, N_2, by heating?

$$NH_4NO_2 \longrightarrow 2\ H_2O + N_2.$$

2. In the preparation of a quantity of nitrogen, N_2, in the laboratory, 34.5 grams of sodium nitrite, $NaNO_2$, was used. How many grams of ammonium chloride, NH_4Cl, would be needed in the reaction?

$$\begin{array}{ll} (a) & NaNO_2 + NH_4Cl \longrightarrow NaCl + NH_4NO_2. \\ (b) & \underline{NH_4NO_2 \qquad\quad \longrightarrow \quad 2\ H_2O + N_2.} \\ \text{Final equation:} & NaNO_2 + NH_4Cl \longrightarrow NaCl + 2\ H_2O + N_2. \end{array}$$

3. What is the volume in liters of the nitrogen, N_2, produced by heating 13.5 grams of ammonium nitrite, NH_4NO_2?

$$NH_4NO_2 \longrightarrow 2\ H_2O + N_2.$$

4. How many liters of nitrogen, N_2, may be produced by heating a mixture of 28 grams of sodium nitrite, $NaNO_2$, with enough ammonium chloride, NH_4Cl, to combine with it completely?

Final equation: $NH_4Cl + NaNO_2 \longrightarrow NaCl + 2\ H_2O + N_2$.

5. How many grams of ammonium nitrite, NH_4NO_2, would be needed to produce, by heating, 400 liters of nitrogen?

$$NH_4NO_2 \longrightarrow 2\ H_2O + N_2.$$

6. How many pounds of sodium nitrite, $NaNO_2$, would be required to react completely with 80 pounds of ammonium chloride, NH_4Cl, in the preparation of nitrogen, N_2?

Final equation: $NH_4Cl + NaNO_2 \longrightarrow NaCl + 2\ H_2O + N_2$.

7. How many liters of nitrogen, N_2, would be produced by burning 500 liters of ammonia, NH_3, in oxygen?

$$4\ NH_3 + 3\ O_2 \longrightarrow 6\ H_2O + 2\ N_2.$$

8. A piece of material was found to be 95% pure sodium nitrate, $NaNO_3$. What per cent of the entire mass is made up of nitrogen?

9. How many grams of water would be produced during the decomposition of 10 grams of ammonium nitrite, NH_4NO_2, in the preparation of pure nitrogen?

$$NH_4NO_2 \longrightarrow 2\ H_2O + N_2.$$

10. How many pounds of ammonium chloride, NH_4Cl, are required to produce 50 liters of nitrogen by its reaction with sodium nitrite, $NaNO_2$, in hot solution?

Final equation: $NH_4Cl + NaNO_2 \longrightarrow NaCl + 2\ H_2O + N_2$.

11. What is the relative density of nitrogen which weighs 1.25 grams per liter, compared with oxygen which has a weight of 1.43 grams per liter?

The Atmosphere

The atmosphere is composed of a mixture of several gases which may be separated by mechanical means. Lack of a definite boiling point and unequal solubility of its constituents in water prove that, while its composition is practically constant, it is not a definite compound.

1. What weight of phosphorus, P, would be needed to remove the oxygen from 600 liters of air? Assume that oxygen forms 23.2% of air by weight.

$$4\ P + 5\ O_2 \longrightarrow 2\ P_2O_5.$$

2. How much nitrogen, of an impure grade, could be obtained by the action of red hot copper, Cu, on air if 10 grams of copper is used up in the reaction? Assume that the air is 23.2% oxygen by weight.

$$2\ Cu + O_2 \longrightarrow 2\ CuO.$$

3. A solution of sodium pyrogallate completely absorbs the oxygen from a sample of air. 89.1 cc. of an air sample is placed in contact with sodium pyrogallate solution and allowed to remain until all of the oxygen is absorbed. The residual gas is measured and found to occupy 70.8 cc. What is the per cent of oxygen in this sample of air?

4. A 95 cc. sample of air is found to contain 20.7% oxygen by volume. After the oxygen is absorbed from the air, the residue of the sample is kept in contact with red hot magnesium until there is no further loss in volume. 0.75 cc. of gas remains after this treatment. What is the per cent of pure nitrogen in the original sample of air?

5. How many grams of copper, Cu, are used up in absorbing the oxygen from a 1500 cc. sample of air, if the copper forms cupric oxide, CuO, and the air contains 20.0% oxygen?

CHAPTER VIII. THE GAS LAWS

Boyle's Law

Of the three forms of matter, gases are most affected in volume by changes in temperature or pressure. This is because gases by their nature are more loosely constructed than the other forms of matter, having relatively much larger distances between the particles making up the mass than do solids or liquids. Gases are composed ultimately of immeasurably minute particles called *molecules* (see page 6), which are constantly in motion, with high velocity, in all directions.

These molecules in gases travel in straight lines, changing their directions only when they collide with other molecules or with surfaces they cannot penetrate. When such collisions with surfaces take place, since the molecules are moving and have mass (that is, are composed of matter), the surfaces are subjected to a constant bombardment which tends to set them in motion. When the surfaces are rigid and fail to move under this bombardment, it is because they are opposing the forces of the molecules with an equal opposite force. Consequently, when a gas is inclosed in a vessel, the gas is confined under pressure, the result of resistance to motion by the walls of the vessel.

If without changing the temperature of the gas, it is forced by some means to occupy only half the volume it originally filled, there will be just twice as many particles striking a given area of surface of the walls of the vessel in a given time. In order to remain in place, the wall must exert twice as much pressure against the gas.

In other words, the pressure has been doubled, when the gas volume is reduced to one half. If we triple the effort of the walls of the vessel against the inclosed gas, the gas is forced to occupy one third of its original volume. This relationship holds good for all ordinary ranges of pressure, so that we may say that *the volume of a gas varies inversely with the pressure exerted upon it.* This is a statement of Boyle's Law, and the following table shows this relationship between pressure and corresponding volume of a given quantity of a gas.

VOLUMES OF OXYGEN AT DIFFERENT PRESSURES, TEMPERATURE REMAINING CONSTANT

QUANTITY OF GAS	VOLUME V	PRESSURE P	$\frac{K}{PV}$
32 grams	179.2 liters	95 mm.	170,240
32 grams	89.6 liters	190 mm.	170,240
32 grams	44.8 liters	380 mm.	170,240
32 grams	22.4 liters	760 mm.	170,240
32 grams	11.2 liters	1,520 mm.	170,240
32 grams	5.6 liters	3,040 mm.	170,240
32 grams	2.8 liters	6,080 mm.	170,240
32 grams	1.4 liters	12,160 mm.	170,240

In this table it will be noticed that as the pressure increases, the volume decreases in exact proportion. The proof of this is in the fact that the product of the pressure upon the gas by the corresponding volume of the gas is a constant quantity. This can be true only when the two variables, pressure and volume, are inversely proportional to each other.

The statement that the volume of a gas varies inversely with the pressure upon it may be put as a proportion:

$$V : V' : : P' : P.$$

Since the product of the means is equal to the product of the extremes, we have:

$$V'P' = VP.$$

Since this is progressively true for changing volumes of gases under changing pressures, we find that these products are equal to a constant value, which may be expressed in an equation as K.

$$V'P' = K = VP,$$

in which, V is the original measured volume,
P is the original measured pressure,
P' is the new pressure, and
V' is the volume corresponding to the new pressure.

This may also be expressed as follows:

$$V' = V \times \frac{P}{P'}.$$

The use of this formula may be illustrated as follows:

Problem: A quantity of carbon monoxide (a gas) is found to have a volume of 300 cc. when measured under a pressure of 730 mm. of mercury. What volume would this gas occupy under standard pressure?

Solution: The formula is as stated:

$$V' = V \times \frac{P}{P'}.$$

Substituting values given in the problem, and canceling,

$$V' = \overset{75}{\cancel{300}}\text{ cc.} \times \frac{\overset{73}{\cancel{730}}}{\underset{\underset{19}{\cancel{76}}}{\cancel{760}}} = \frac{5475}{19} = 288.1\text{ cc.}$$

The new volume of the gas, under standard pressure, 760 mm. of mercury, is thus found to be 288.1 cc. of carbon monoxide.

Charles' Law

Heat (energy) applied to a gas causes the molecules to move more rapidly, since it is energy alone that gives motion to molecules. Moving more rapidly, the molecules batter with greater force against the restraining walls of a vessel confining the gas. If the walls are immovable, the gas is at once under increased pressure. If, however, the walls of the vessel are movable and oppose the gas particles with a uniform resistance, then the walls of the vessel will be forced outward until the total pressure of the molecules battering against a unit of surface of the walls becomes the same as the pressure at the beginning of the heat change. Briefly stated, the gas expands when heated and kept under unchanging pressure.

By careful experiment it has been learned that if a certain volume of a gas is measured at 0° C., and then subjected to a change in temperature, the volume changes $\frac{1}{273}$ of the volume at zero for each degree change in temperature.

Using the gas helium as an example, we find that under unchanging pressure the following ideal volumes will correspond to the temperatures given:

273 cc.	0° C.	273 abs.
272 cc.	−1° C.	272 abs.
271 cc.	−2° C.	271 abs.
270 cc.	−3° C.	270 abs.
170 cc.	−103° C.	170 abs.
70 cc.	−203° C.	70 abs.
50 cc.	−223° C.	50 abs.
10 cc.	−263° C.	10 abs
5 cc.	−268° C.	5 abs.

Helium liquefies at −269° C., under 760 mm. pressure.

The contraction of the gas is due not to a contraction in the volume of the molecules themselves, but of the intermolecular spaces. Disregarding the space occupied by the molecules themselves, the intermolecular spaces would be reduced to zero and the gas incapable of further decrease in volume at −273° C., theoretically, provided the substance had remained in the gas state until that temperature was reached.

This condition could only occur in the entire absence of energy, otherwise the molecules would have some motion and would necessarily be at some distance from each other. The entire absence of heat indicated by a temperature of −273° C., is called *absolute zero.* Below this temperature it would be impossible to go.

If now we start from this absolute zero and count upward, using degrees of the same magnitude as those of the Centigrade scale, −200° C. becomes 73° abs., −100° C. becomes 173° abs., and 0° C. becomes 273° abs. In other words, to convert temperature expressed in Centigrade degrees to temperature expressed in absolute degrees, we add 273 algebraically to the Centigrade reading.

Referring now to the third column in the table on page 32, we find that the volumes of the gas correspond exactly to the absolute temperature; that is, *the volume of a gas varies directly with the absolute temperature of the gas, provided the pressure remains constant.* This is expressed as a formula thus: $V' : V : : T' : T$ or $V'T = VT'$. Therefore, $V' = V \times \frac{T'}{T}$.

Example: 500 cc. of a gas is measured at 18° C. and 740 mm. Later the temperature rises to 32° C. What volume will the gas now occupy, pressure remaining constant?

Add 273 to the temperature readings:

18° C.+273 = 291° abs.
32° C.+273 = 305° abs.

$$500 \times \frac{305}{291} = \frac{152{,}500}{291} = 524.05 \text{ cc., corrected volume.}$$

To combine the effects of change of pressure and change of temperature, acting simultaneously, we multiply the original volume of the gas by the pressure correction factor, $\frac{P}{P'}$, and the temperature correction factor, $\frac{T'}{T}$, in succession: in which

P represents original pressure,
P' represents final pressure,
T represents original temperature, absolute,
T' represents final temperature, absolute.

As an example, 300 cc. of argon is measured at a temperature of 27° C. and 740 mm. of mercury pressure. What volume will this gas occupy at $-23°$ C. and 760 mm. of mercury pressure? The formula used is:

$$V' = V \times \frac{P}{P'} \times \frac{T'}{T}.$$

Substituting values in the formula,

$$V' = 300 \text{ cc.} \times \frac{740}{760} \times \frac{273 + (-23)}{273 + 27},$$

or $$V' = \overset{1}{\cancel{300}} \text{ cc.} \times \frac{\overset{37}{\cancel{740}}}{\underset{\underset{19}{\cancel{38}}}{\cancel{760}}} \times \frac{\overset{125}{\cancel{250}}}{\underset{1}{\cancel{300}}} = \frac{4625}{19} = 243.4 \text{ cc. argon.}$$

Correction for Tension of Aqueous Vapor

Water evaporates at all temperatures, the amount of water remaining in a gas in contact with water depending upon the temperature of the gas. Since the water vapor makes up part of the gas measured, and since it is evenly distributed throughout the mass of the gas, the simplest way to make correction for the water vapor present in the gas is through the pressure exerted by the water vapor.

In a mixture of gases, each gas occupies the entire volume of the vessel, but exerts only a part of the total pressure of the gas within the vessel, the partial pressures of the gases being proportional to the amounts of the gases present.

It is a comparatively simple matter to determine the pressure exerted by water vapor at any temperature. After this has been done, the results of such determinations may be compiled into a table indicating the pressure of aqueous vapor throughout the ordinary range of temperatures. To determine the true pressure of the gas alone, freed from water vapor, it is only necessary to subtract from the total pressure of the gas including the water vapor, the pressure due to the water vapor alone, as shown in the table, for the temperature at which the gas is measured. The result is the pressure the gas alone would exert occupying the same volume as measured.

To make correction for the aqueous vapor in a gas collected over water, all that is necessary is to subtract from the measured pressure as indicated usually by the barometer, the tension (pressure) of aqueous vapor at the temperature at which the gas was measured.

The complete formula for the correction of gases to new conditions, including change of temperature, change of pressure, and correction for the presence of aqueous vapor is as follows:

$V' = V \times \frac{P-a}{P'} \times \frac{T'}{T}$, in which a stands for the pressure due to tension of aqueous vapor at the temperature of the gas.

Example: A quantity of argon collected over water is found to have a volume of 800 cc., with the temperature at 20° C. and the pressure 740 mm. of mercury. What volume would this gas occupy at standard conditions? The formula is:

$$V' = V \times \frac{P-a}{P'} \times \frac{T'}{T}.$$

Aqueous tension is found, by referring to the table, to be 17.36 mm. of mercury at 20° C.

$$V' = 800 \times \frac{740-17.36}{760} \times \frac{273}{293},$$

or
$$V' = \overset{20}{\cancel{800}} \times \frac{722.64}{\underset{19}{\cancel{760}}} \times \frac{273}{293} = 708.75 \text{ cc. of argon.}$$

Correction for Difference of Level

When the volume of a gas inclosed in a tube over a liquid is measured, it often happens that the surface of the liquid inside the tube is higher or lower than that outside it. Of course, in such a case the pressure on the inclosed gas is different from that of the air.

In case the liquid inside the tube is higher than that outside, it will be seen that the air pressure is balanced by the pressure of the inclosed gas plus the pressure due to the column of liquid, measured vertically from the level of the surface outside to the level of the surface inside the tube. To obtain the true pressure of the gas in this case, it is necessary to subtract the pressure of the liquid column (expressed in millimeters of mercury) from the reading of the barometer. The difference is the pressure of the gas alone.

Example: 350 cc. of a gas is measured in a tube over mercury, the level of the mercury inside the tube being 27 mm. higher than that outside. The barometer reads 748 mm. What is the true pressure of the gas?

Solution: Subtract 27 mm. from 748 mm. The difference, 721 mm., is the true pressure of the gas.

In case the mercury level outside the tube containing the inclosed gas is higher than that inside, it is evident that the gas is supporting the pressure of the air plus the pressure of the column of mercury due to the difference of level.

In this case, to find the true pressure of the gas we must add the pressure of the column of mercury measured from the level outside to the level inside the tube.

Example: 600 cc. of nitrogen is inclosed in a tube over mercury. When measured, the mercury outside the tube is 65 mm. higher than that inside. What is the true pressure of the gas if the barometer reads 740 mm.?

Solution: Add 65 mm. to the barometer reading, 740 mm. The result, 805 mm., is the true pressure of the gas.

1. Convert the following thermometer readings to the absolute scale: 25° C.; −73° C.; 68° F.; 127° C.; −40° F.

2. A quantity of hydrogen gas measured 300 cc. at 27° C. What volume would it occupy at −73° C., pressure upon the gas remaining constant?

3. A quantity of nitrogen measures 500 cc. at a pressure of 720 mm. of mercury. What would be its volume at a pressure of 750 mm. of mercury, the temperature of the gas remaining constant?

4. At a temperature of −23° C. and a pressure of 750 mm. of mercury, a gas was found to measure 280 cc. At a later reading the temperature was found to be 27° C. and the pressure had changed to 560 mm. of mercury. What was the new volume of the gas?

5. Make the following temperature conversions:

12° F. to degrees C.	16° C. to degrees F.
132° F. to degrees absolute	−85° F. to degrees C.

212° absolute to degrees F.

6. A quantity of oxygen measures 82 cc. at 16° C. At what temperature would it measure 95 cc., the pressure upon the gas remaining constant?

7. A quantity of carbon dioxide is measured and the volume found to be 65.2 cc. when the temperature is 20° C. and the pressure is 740 mm. of mercury. What volume would be occupied by this gas at standard temperature and pressure?

8. A quantity of hydrogen was inclosed in a eudiometer over mercury and the volume determined to be 48 cc. The pressure was found to be 730 mm. of mercury and the temperature 98° C. The gas was later measured again, the volume this time being 49 cc. The temperature was still 98° C. What was the pressure in millimeters of mercury at the time of the second measurement?

9. An open vessel was found to contain 2 grams of air. How many grams of air would the vessel contain if the pressure were doubled?

10. At what pressure would one liter of hydrogen weigh two grams, temperature remaining at 0° C.?

11. At what temperature, pressure remaining at 760 mm. of mercury, would one liter of oxygen weigh four grams?

12. A certain gas was collected over mercury and was found to measure 600 cc., at a temperature of 25° C. and a pressure of 738 mm. of mercury. What volume would this gas occupy if measured over water under the same conditions of temperature and pressure?

13. A gas is collected over mercury and the following noted: volume of gas 580 cc., temperature 26° C., barometric reading 758 mm. of mercury, height of mercury column inside of measuring tube 18 mm. above outside level of mercury. Find the volume this gas would occupy under standard conditions of temperature and pressure.

14. A quantity of hydrogen is collected over mercury and the following data secured: volume of gas 348 cc., temperature of gas 22° C., barometer reading 745 mm. of mercury, mercury column inside of measuring tube 24 mm. above the level of the mercury outside. Find the volume this gas would occupy at the same temperature and pressure if it were collected over water, inside and outside levels being the same.

15. An inflated rubber bag contains 500 cc. of air at an atmospheric pressure of 14.7 lbs. per square inch. What would be its volume if it were to be immersed in water to a depth of 40 feet? Fresh water weighs 62.4 pounds to the cubic foot.

16. A balloon is partially inflated at sea level with 500,000 cubic feet of hydrogen. What volume would this hydrogen occupy at such a height that the air pressure would sustain a column of mercury 650 mm. high?

17. A glass globe is found to contain 3.1 grams of air at a pressure of 740 mm. of mercury. At what pressure will the same vessel contain 3.9 grams of air, temperature remaining constant?

18. One liter of nitrogen weighs 1.25 grams at standard temperature and pressure. What pressure is exerted by 3.7

grams of nitrogen confined in the same vessel, temperature remaining unchanged?

19. A quantity of argon is found to have a volume of 150.4 cc. under standard temperature and pressure. Later the temperature is found to be 57° C. while the pressure is the same as before. What is the volume of the argon under the new conditions?

20. 70 cc. of helium is measured at 80° C. If the pressure does not vary, what will the volume of this helium become at 310° C.?

21. A quantity of carbon dioxide is collected over mercury and found to have a volume of 97.8 cc. The pressure of the gas is 738 mm. of mercury and the temperature is 18° C. Correct the measured volume to standard temperature and pressure.

22. A quantity of air is inclosed in a eudiometer over mercury and the following data noted:

Volume of air	51.3 cc.
Temperature of air	21.0° C.
Barometer reading	758.0 mm. of mercury.
Height of mercury inside tube above that outside of tube	110.0 mm.

What volume would this air occupy under standard temperature and pressure?

23. The oxygen produced by the decomposition of 0.3 grams of potassium chlorate is collected by the displacement of water and measured at a temperature of 20° C. and a pressure of 750 mm. of mercury. Assuming that 1 cc. of the oxygen is dissolved by the water, what volume does the gas occupy?

24. At what temperature will 350 cc. of a gas, measured at 21° C., occupy two thirds of that volume, pressure remaining unchanged?

25. At what temperature will 600 cc. of a gas occupy twice as much space, pressure remaining unchanged?

26. At what temperature will a certain amount of gas occupy 950 cc., if it now occupies a volume of 390 cc. at 16° C.?

27. An increase in temperature of how many degrees C. is required to double the volume of 3 grams of a certain gas, if the original measurement is made at 15° C.?

28. How many degrees C. rise in temperature is required to double the volume of 3 grams of a certain gas, if the original measurement is made at 20° C.?

29. Convert to the Centigrade scale: 20° F., 32° F., 96° F., 68° F.

30. Convert to the Fahrenheit scale: 16° C., 18° C., 20° C., 0° C.

31. Convert to the absolute scale: 90° C., 96° F., 30° C., 68° F., 320° F., 78° F., 60° C., 18° C.

32. A certain amount of oxygen, collected over water, is found to occupy 320 cc. at a temperature of 18° C. and at a standard pressure. What volume would this oxygen occupy under the same conditions but with the water vapor removed?

33. Oxygen, measured at standard temperature and pressure, has a relative density of 1.1 compared with air at standard temperature and pressure. What will be the relative density of oxygen at 30° C. compared with air measured at standard temperature and pressure?

34. A certain amount of nitrogen, collected over water, occupies 1000 cc. at a pressure of 760 mm. Upon analysis, it is found that the gas is 3.1% water vapor by volume. At what temperature was the gas measured?

35. Helium at standard temperature and pressure has a specific gravity of 0.1188. What is the specific gravity of helium at 96° C. compared with air at standard temperature and pressure?

36. A certain amount of hydrogen, collected over water, is found to occupy 650 cc. at a temperature of 22° C. and at standard pressure. What part of the total volume would the hydrogen alone occupy if measured under the same conditions?

37. Nitrogen, measured at standard temperature and pressure, has a relative density of 0.96, compared with air at standard temperature and pressure. What is the relative density of nitrogen at 24° C. and a pressure of 760 mm. of mercury, compared with air under the same conditions?

38. A certain amount of oxygen, collected over water, was found to have a volume of 63 cc. under a pressure of 760 mm. of mercury. By analysis it was found that the water vapor present would occupy 2.61 cc. if measured alone under the same conditions. What was the temperature at which the gas was measured?

39. Hydrogen, measured at standard temperature and pressure, has a relative density of 0.07 compared with air under the same conditions. What is the relative density of hydrogen, measured at 70° C. and a pressure of 760 mm. of mercury, compared with air at standard temperature and pressure?

CHAPTER IX. SOLUTIONS AND NEUTRALIZATION

Solutions

A solution is a condition in which one or more substances have so completely mingled with another substance, without undergoing chemical change, that their particles cannot be distinguished by any means at our command. The most common type of solution is that of a solid in a liquid. Another common type is that of a gas in a liquid. Examples are: for the solid in a liquid, common salt in water; for the gas in a liquid, nitrogen in water.

The amount of substance (solute) which may be dissolved in a given amount of a liquid (solvent) depends upon several factors. In the case of a solid (solute) dissolved in a liquid (solvent) the amount dissolved depends upon: (*a*) the nature of the solute; (*b*) the nature of the solvent; and (*c*) the temperature. Usually, greater amounts of solids dissolve in liquids as the temperature rises.

In the case of gases, the amount of a gas which will dissolve in a given amount of a certain liquid depends upon: (*a*) the nature of the liquid (solvent); (*b*) the nature of the gas (solute); (*c*) the temperature; and (*d*) the pressure. With regard to pressure, the amount of gas soluble in a liquid is proportional to the pressure. This last statement is true for the ordinary range of pressures.

1. At a temperature of 0° C. and a pressure of 760 mm. of mercury, 1713 cc. of carbon dioxide will dissolve in one liter of water. How much carbon dioxide gas would escape as bubbles from a 500 cc. glass of soda water at 0° C., if it is drawn from a tank in which the water is saturated with carbon dioxide at a pressure of 1000 mm. of mercury?

2. 32.5 grams of zinc reacts completely with sulfuric acid, H_2SO_4, and the solution is evaporated to dryness. What solid is left and how much does it weigh?

3. At 40° C. and a pressure of 760 mm. of mercury, 1 gram of water will dissolve 20.4 cc. of sulfur dioxide. How many cubic centimeters of sulfur dioxide will dissolve in 250 cc. of water at a pressure of 738.6 mm. of mercury, if the temperature is 40° C.?

4. What volume in liters of ammonia, NH_3, may be produced by boiling 100 grams of a solution of ammonium hydroxide, 20% NH_4OH? The equation is:

$$NH_4OH \longrightarrow H_2O + NH_3.$$

Neutralization

When hydrogen chloride is dissolved in water a part of its molecules dissociate or break up to form *ions*. Each HCl molecule so dissociated yields one H^+ ion bearing one positive charge of electricity and one Cl^- ion bearing one negative charge of electricity. The change in the hydrogen chloride, HCl, is represented by the equation:

$$HCl \longrightarrow H^+ + Cl^-.$$

These ions wander freely about through the solution (*ion* means *wanderer*) and give to the solution its chemical properties. The acid properties of the solution are due to the presence of the hydrogen, H^+, ions. This is the reason for the great chemical difference between hydrogen chloride and hydrochloric acid, which is a water solution of the gas. The dry gas hydrogen chloride is entirely neutral in its action upon litmus paper which is free from moisture. This gas exhibits none of the acid properties of the water solution.

We can therefore see that the acid properties of hydrochloric acid are due to the presence of the hydrogen, H^+, ions and not to the presence of hydrogen, H. The Cl^- ions in the solution are also ready to enter into a chemical reaction more readily than would be the atoms of chlorine, Cl. As these H^+ ions and Cl^- ions wander about they frequently collide and again form molecules of hydrogen chloride, HCl, the single positive electrical charge on each hydrogen, H^+, ion being neutralized by the single negative electrical charge on the chlorine, Cl^-, ion.

At the same time more of the hydrogen chloride, HCl, molecules dissociate to form ions, thus keeping the ion concentration continually the same. The percentage of dissociated molecules for any ionizable substance varies with the nature of the substance and with the extent to which the solution is diluted, but is never up to 100% in any case. The dissociation of the hydrogen sulfate, H_2SO_4, molecule in sulfuric acid is represented in the equation:

$$H_2SO_4 \longrightarrow H^+ + H^+ + SO_4^{--}.$$

In this case we have the formation of two H^+ ions and one SO_4^{--} ion, from each dissociated hydrogen sulfate molecule, the two negative charges of electricity on the sulfate, SO_4^{--}, ion just offsetting the two single positive charges on the two hydrogen, H^+, ions. In nitric acid the hydrogen nitrate, HNO_3, molecules which are dissociated break up into H^+ ions and NO_3^- ions according to the equation:

$$HNO_3 \longrightarrow H^+ + NO_3^-.$$

We can thus see that when an acid is dissolved in water a part of its molecules are dissociated into ions and one of the ions in each case is H^+. It is this H^+ ion to which any acid owes its acid properties.

When sodium hydroxide, NaOH, is dissolved in water a part of the NaOH molecules dissociate to form sodium, Na^+, ions and hydroxyl, OH^-, ions. The equation is:

$$NaOH \longrightarrow Na^+ + OH^-.$$

Potassium hydroxide, KOH, molecules in the same way form potassium, K^+, ions and hydroxyl, OH^-, ions, according to the equation:

$$KOH \longrightarrow K^+ + OH^-.$$

Ammonium hydroxide, NH_4OH, molecules form, when dissociated, ammonium, $(NH_4)^+$, ions and hydroxyl, OH^-, ions as shown by the equation:

$$NH_4OH \longrightarrow (NH_4)^+ + OH^-.$$

Each base that ionizes or dissociates forms, as one of the ions, the hydroxyl ion, OH^-. It is this OH^- ion to which the base owes its basic properties, just as an acid owes its acid properties to the H^+ ions present in its solution.

Now suppose that we mix a solution of sodium hydroxide, NaOH, with a solution of hydrogen chloride (hydrochloric acid), HCl. In such a mixture there would be present Na^+ ions and OH^- ions from the base and H^+ ions and Cl^- ions from the acid. We might represent this by the ionic equation:

$$NaOH + HCl \longrightarrow Na^+ + OH^- + H^+ + Cl^-.$$

These four kinds of ions are present in the solution at the same time and are constantly moving about and colliding with each other. From these collisions new molecules result and these molecules in turn dissociate to form the same ions. Also some of the original molecules of NaOH and HCl would always be present, since the dissociation is never absolute.

We see, therefore, that the condition in such a solution is far from static and that a state of dynamic equilibrium is maintained by the constant dissociation of molecules and the subsequent recombination of the resulting ions. In such a solution there would be present at the same instant the following:

$$H^+ + Cl^- + Na^+ + OH^- + NaOH + HCl + NaCl + H_2O.$$

In this seeming condition of equilibrium, however, there is one molecule, H_2O, which is dissociated to such a slight extent that it is to all intents removed from the field of action as rapidly as formed. This, then, would result in using up one H^+ from the acid for each OH^- of the base to form one molecule of undissociated H_2O.

If we use, then, the proper amounts of the base and the acid, we should have a resulting solution which would have the properties of neither base nor acid but would be neutral. The process of mixing a base and an acid to form a neutral solution is called *neutralization.* We may now rewrite the equation for the neutralization of hydrochloric acid with sodium hydroxide as follows:

$$NaOH + HCl \longrightarrow Na^+ + OH^- + H^+ + Cl^- \longrightarrow H_2O + Na^+ + Cl^-$$

Since the NaCl molecule formed from the Na^+ and Cl^- ions dissociates so readily in solution, the salt may be obtained from the solution only by concentration. NaCl is thus obtained by forcing the Na^+ and Cl^- ions to form undissociated molecules of NaCl by the equation:

$$Na^+ + Cl^- \longrightarrow NaCl.$$

The neutralization of sulfuric acid with potassium hydroxide is as follows:

$$2\,KOH + H_2SO_4 \longrightarrow K^+ + K^+ + OH^- + OH^- + H^+ + H^+ + SO_4^{--} \longrightarrow 2\,H_2O + K^+ + K^+ + SO_4^{--}.$$

In this case the K_2SO_4 may be obtained by evaporation of the water from the solution. In every case of the neutralization of an acid with a base the H^+ of the acid combines with the OH^- of the base to form an undissociated molecule of H_2O. The metallic ion of the base may then be forced to combine with the negative ion of the acid to form an undissociated molecule of a salt by concentration of the solution.

The calculation of the quantities of materials involved in neutralization may be performed as with other chemical reactions.

1. Write equations, indicating ions, for the neutralization of sodium hydroxide with nitric acid; for the neutralization of hydrochloric acid with potassium hydroxide.

2. Write equations, indicating ions, for the neutralization of potassium hydroxide with sulfuric acid; for the neutralization of calcium hydroxide with acetic acid.

3. What weight in grams of sodium hydroxide, NaOH, would be necessary to neutralize 98 grams of sulfuric acid?

4. How many grams of calcium hydroxide, $Ca(OH)_2$, would be needed to neutralize 126 grams of nitric acid, HNO_3?

5. What weight in grams of hydrogen chloride, HCl, would be needed to neutralize 35 grams of concentrated ammonium hydroxide, 50% NH_4OH?

6. A solution of sodium hydroxide, NaOH, was exactly neutralized by hydrochloric acid, HCl, and the resulting solution evaporated to dryness. The salt residue was found to weigh 117 grams. What weights in grams of base and of acid were used?

7. How many grams of water would be formed by the neutralization of 80 grams of sodium hydroxide, $NaOH$, with phosphoric acid, H_3PO_4?

8. How many grams of potassium hydroxide, KOH, would be formed by the action of 78 grams of potassium, K, on water? If the resulting solution were neutralized with nitric acid, HNO_3, and then evaporated to dryness, how many grams of potassium nitrate, KNO_3, would be left as residue?

9. A solution of potassium hydroxide is prepared with 10 grams of KOH per liter of solution. How many cubic centimeters of this solution will be required to neutralize 50 cc. of a solution of sulfuric acid, containing 10 grams of H_2SO_4 per liter of solution?

10. At 15° C. a maximum of 0.137 gram of calcium hydroxide, $Ca(OH)_2$, will dissolve in 100 grams of water. How many grams of hydrochloric acid, 25% HCl, will be required to neutralize 40 grams of a solution of calcium hydroxide saturated at 15° C.?

11. 20 grams of a 4% solution of sodium hydroxide, $NaOH$, is required to neutralize 30 grams of vinegar (dilute acetic acid, $HC_2H_3O_2$). What is the per cent acid strength of the vinegar?

12. What volume in liters of hydrogen chloride, HCl, is required to combine exactly with 1.4 liters of ammonia, NH_3?

13. 300 liters of sulfuric acid solution, containing 30 grams of H_2SO_4 per liter of solution, is neutralized by adding anhydrous sodium carbonate, Na_2CO_3. How many grams of sodium carbonate would be required for the reaction?

14. 40 cc. of sodium hydroxide, made up with 20 grams of NaOH per liter of solution, is neutralized with a solution of oxalic acid, containing 9 grams of $H_2C_2O_4$ to the liter of solution. How many liters of oxalic acid solution would be needed for the reaction?

15. 100 grams of a 10% solution of nitric acid, HNO_3, is neutralized by a 30% solution of potassium hydroxide, KOH. How many grams of the base solution are required for the neutralization?

16. If, in neutralizing a certain amount of nitric acid, HNO_3, with sodium hydroxide, NaOH, 12.5 cc. of water is formed, what weights of pure nitric acid and of sodium hydroxide have been employed?

17. How many grams of hydrogen chloride, HCl, would be needed to exactly neutralize 80 cc. of a potassium hydroxide solution containing 35 grams of KOH per liter of solution?

18. How many grams of a 5% solution of sulfuric acid, H_2SO_4, would be required to exactly neutralize 40 grams of a 10% solution of sodium hydroxide, NaOH?

19. How many grams of pure potassium hydroxide, KOH, would be needed to exactly neutralize the acetic acid, $HC_2H_3O_2$, in 5 kilograms of vinegar if the acid forms 6% of the solution by weight?

20. 3 grams of sodium bicarbonate, $NaHCO_3$, are dissolved in water, and dilute hydrochloric acid, 20% HCl, is added until the reaction is complete. How many grams of the acid solution are required?

21. 50 grams of a solution containing lithium hydroxide, LiOH, is neutralized with a solution of nitric acid containing 10 grams of HNO_3 per 10 cc. of solution. It requires 40 cc. of the acid to effect neutralization. What is the strength of lithium hydroxide in the sample tested?

22. 25 grams of a solution of sulfuric acid is neutralized with 60 grams of a solution of potassium hydroxide containing 60 grams of KOH per 1000 grams of solution. What is the strength of the acid solution?

23. Sulfuric acid, 98% H_2SO_4, has a specific gravity of 1.84. How many cubic centimeters of this sulfuric acid are required to neutralize the potassium hydroxide contained in 130 cc. of a solution made with 15 grams of KOH per 100 cc. of solution?

24. Hydrochloric acid solution with a specific gravity of 1.125 contains 24.78% HCl. How many cubic centimeters of such a solution are required to neutralize 92.5 cc. of a sodium hydroxide solution containing 40 grams of NaOH per liter of solution?

25. Nitric acid with a specific gravity of 1.125 contains 21% HNO_3. How many cubic centimeters of this acid are required to neutralize 39.2 cc. of a solution of potassium hydroxide containing 56 grams of KOH per liter of solution?

CHAPTER X. SALTS

A salt may be defined as a compound of a metal with the characteristic radical of an acid. A salt may be prepared in any one of five ways:

1. By direct combination of elements.

$$Cu + Cl_2 \longrightarrow CuCl_2.$$

2. By substitution of the hydrogen of an acid by a metal.

$$Zn + 2\,HCl \longrightarrow ZnCl_2 + H_2.$$
$$ZnO + 2\,HCl \longrightarrow ZnCl_2 + H_2O.$$

3. By substitution of one metal for another in a salt.

$$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu.$$

4. By reaction between two salts to produce two other salts.

$$MgCl_2 + 2\,AgNO_3 \longrightarrow Mg(NO_3)_2 + 2\,AgCl.$$

5. By reaction between an acid and a salt to produce another acid and another salt.

$$NaCl + H_2SO_4 \longrightarrow NaHSO_4 + HCl.$$
$$2\,NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2\,HCl.$$

A normal salt is one that contains no displaceable hydrogen in its molecule.

An acid salt is one that contains displaceable hydrogen in its molecule.

Examples of normal salts are: cupric chloride, $CuCl_2$; zinc chloride, $ZnCl_2$; ferrous sulfate, $FeSO_4$; magnesium nitrate, $Mg(NO_3)_2$; and sodium sulfate, Na_2SO_4.

Examples of acid salts are: sodium acid sulfate, $NaHSO_4$; sodium acid carbonate, $NaHCO_3$; and calcium acid carbonate, $CaH_2(CO_3)_2$.

1. How many grams of ammonium sulfate, $(NH_4)_2SO_4$, would result from the neutralization of 94 grams of sulfuric acid with ammonium hydroxide?

2. How many pounds of nitric acid can be made from 100 pounds of sodium nitrate if the salt is only 85% $NaNO_3$?

3. It requires 50 cc. of a sodium hydroxide solution to neutralize a 100 cc. sample of vinegar. The sodium hydroxide solution contains 40 grams of NaOH in one liter of solution. What is the acetic acid, $HC_2H_3O_2$, strength of the vinegar?

4. A solution of potassium hydroxide containing 5.6 grams of KOH in 100 cc. of solution was neutralized by a solution of sulfuric acid containing 17.6 grams of H_2SO_4 in 100 cc. of solution. How many cubic centimeters of the acid would be required to neutralize 500 cc. of the base solution?

5. What weight of crystallized calcium nitrate, $Ca(NO_3)_2 \cdot 4\ H_2O$, will be produced in neutralizing 20 grams of calcium hydroxide, $Ca(OH)_2$, with dilute nitric acid, HNO_3?

6. How many grams of copper nitrate, $Cu(NO_3)_2$, would be required to produce 15 grams of cupric hydroxide, $Cu(OH)_2$, when reacted upon by potassium hydroxide solution?

7. Sodium hydroxide is often prepared by boiling a mixture of milk of lime, $Ca(OH)_2$, and sodium carbonate, Na_2CO_3, solution. Provided that 85% of the sodium hydroxide produced is recovered in the operation, how much sodium carbonate is required to yield 100 pounds of sodium hydroxide?

8. 100 cc. of a solution of sulfuric acid is mixed with a solution of barium hydroxide, $Ba(OH)_2$, until no further precipitate of barium sulfate, $BaSO_4$, is formed. The solution is filtered,

and the barium sulfate dried and weighed. The weight is found to be 0.315 gram. How many grams of sulfuric acid per 100 cc. are present in the original solution?

9. What weight of potassium sulfate, K_2SO_4, would be formed by neutralizing 600 cc. of sulfuric acid, 40 grams of H_2SO_4 to the liter, with potassium hydroxide solution?

CHAPTER XI. VALENCE

Valence is the measure of the number of atoms of hydrogen which can be held in chemical combination by one atom of another element, or which may be chemically displaced by one atom of another element. Since one atom of hydrogen never combines chemically with more than one atom of another element, the hydrogen atom is taken as the standard unit for valence, and to it is assigned the valence 1.

One atom of chlorine combines chemically with one atom of hydrogen to form hydrogen chloride, HCl. Therefore the valence of chlorine is equal to that of hydrogen, that is, 1. One atom of aluminum combines chemically with three atoms of chlorine, or displaces three atoms of hydrogen in a compound. Therefore the valence of aluminum is 3.

In some cases, the valence of an element varies according to the circumstances. Iron has the valences 2 and 3. The valence 2 for iron is characteristic of certain conditions present when the compound is formed. Under certain other conditions, essentially different, iron enters into combination with elements with the valence 3.

An element tends to retain the same valence throughout a series of chemical changes. However, if oxidizing agents are brought in contact with compounds containing the element with variable valence, the tendency is for the element to take the higher valence. If reducing agents are brought in contact with the compound, the tendency is for the element to take the lower valence.

The valence of an element may be determined by comparing it indirectly with hydrogen, the standard. For instance, in the compound aluminum oxide, Al_2O_3, two atoms of aluminum are chemically combined with three atoms of oxygen. The valence of one atom of oxygen is determined from its hydrogen compound, water, H_2O, which gives oxygen a valence of 2.

Now since in aluminum oxide we have three atoms of oxygen, the combined valence of the oxygen atoms is 3×2, or 6. Two atoms of aluminum are combined with oxygen, of a total valence of 6; accordingly each atom of aluminum will have one half the total valence, or 3. The valence of aluminum is thus established indirectly to be 3. Notice that this valence agrees with that previously determined from the compound aluminum chloride, in which the metal was directly substituted for hydrogen.

Radicals, that is, groups of elements which enter into chemical action as units, such as (NO_3), (OH), (SO_4), (CO_3), have definite valences, determined in the same way as the valences of the elements.

In compounds containing more than two elements, we cannot certainly calculate the valences of the individual elements according to the plan here outlined, because the formula of such a compound does not indicate the manner in which the elements are linked together.

1. Give the valence of the underscored element or radical:

$H\underline{Cl}$, $\underline{Na}Cl$, $Na\underline{OH}$, $\underline{Ca}Cl_2$, $Na_2\underline{SO_4}$, $\underline{Zn}SO_4$, $Ca_3(\underline{PO_4})_2$.

2. Give the valence of the underscored elements or radicals in the following compounds:

$H\underline{Cl}$, $\underline{Pb}Cl_2$, $Na\underline{NO_3}$, $\underline{Ag}NO_3$, $\underline{K}OH$, $\underline{Mg}Cl_2$, $K_2\underline{S}$, $\underline{Cu}(NO_3)_2$.

3. Give the valence of the underscored elements or radicals in the following compounds:

$\underline{Al}(OH)_3$, $\underline{Cu}O$, $\underline{C}O_2$, $\underline{N}_2O_3$, $\underline{Ba}SO_4$, $Na\underline{I}$, $\underline{Co}(NO_3)_2$, $\underline{Fe}_2(SO_4)_3$, $Ca\underline{F}_2$, $K_2\underline{CO_3}$.

CHAPTER XII. COMPOUNDS OF NITROGEN

Ammonia and Nitric Acid

The compounds of nitrogen, as was stated in Chapter VII, are unstable. They have been built up slowly in nature and are being consumed faster than nature produces them.

1. How many grams of pure calcium hydroxide, $Ca(OH)_2$, are required to produce 44.1 liters of ammonia, NH_3, by its action on ammonium chloride, NH_4Cl?

2. How many liters of ammonia, NH_3, can be prepared by the action of a base on 3 kilos of sal ammoniac, NH_4Cl?

3. How many grams of ammonium hydroxide, NH_4OH, having a density of 0.90 compared with water, can be made from 500 liters of ammonia? Ammonium hydroxide, specific gravity 0.90 at 15°C., contains 28.32% NH_3.

4. How much nitrogen by volume is needed to produce 1 kilogram of ammonia, NH_3, by the Haber[1] process?

5. How much nitric acid, HNO_3, by weight could be made by the action of sulfuric acid upon 1 kilogram of sodium nitrate, $NaNO_3$?

6. What weight of commercial nitric acid, 70% HNO_3, could be manufactured from one ton of sulfuric acid, 98% H_2SO_4?

7. What volume in liters of ammonia, NH_3, may be produced by the union of 3.2 liters of hydrogen with nitrogen?

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[1] For a detailed description of this process, see page 133.

8. On the assumption that 3% of the nitrogen of the air is converted into nitric acid, what volume of air in cubic feet must be passed through the apparatus to produce 7 pounds of concentrated nitric acid, 70% HNO_3? Consider air as 78.06% nitrogen.

9. How many grams of nitric acid, 70% HNO_3, are required to react with 15 grams of tin, Sn, to form the insoluble, white meta-stannic acid, H_2SnO_3?

10. Copper nitrate, $Cu(NO_3)_2$, decomposes when heated and forms, among other products, black copper oxide, CuO. What weight of crystallized cupric nitrate, $Cu(NO_3)_2 \cdot 6\,H_2O$, is required to produce in this way 1.6 grams of copper oxide?

11. What weight of sulfuric acid, 98% H_2SO_4, is required to manufacture 1 ton of nitric acid, 70% HNO_3?

12. How much water must be added to change one kilogram of 70% nitric acid to a 30% solution?

13. How many liters of ammonia, NH_3, could be formed from ammonium chloride, NH_4Cl, by the action of 22 pounds of sodium hydroxide upon it?

14. How many grams of ammonium chloride, NH_4Cl, would be needed in the production of 400 liters of ammonia, NH_3?

15. How much ammonium hydroxide, NH_4OH, by weight can be made from 300 kilograms of ammonium chloride, NH_4Cl, by the action of calcium hydroxide upon it? Assume that 700 volumes of ammonia can be dissolved in one volume of water at ordinary room temperature.

16. How many liters of ammonia can be prepared from a mixture of hydrogen with 400 liters of nitrogen by the Haber[1] process?

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[1] A detailed description of this process will be found on page 133.

17. How many liters of ammonia can be prepared from 1 ton of ammonium sulfate, if it is 90% $(NH_4)_2SO_4$, by the action of calcium hydroxide upon it?

18. What weight of nitric acid, 70% HNO_3, can be manufactured from 50 tons of sodium nitrate, 95% $NaNO_3$?

19. How many tons of nitric acid, 68% HNO_3, can be manufactured from sodium nitrate, $NaNO_3$, if an excess of the nitrate is heated strongly with 40 tons of sulfuric acid, 98% H_2SO_4?

20. How many tons of sodium nitrate, 99% $NaNO_3$, would be required to make 100 tons of nitric acid, 68% HNO_3?

21. What volume of nitrogen would be used up in making 300 grams of nitric acid, 50% HNO_3, by intensely heating air in an electric arc and combining the oxides of nitrogen formed with water?

22. What weight of salt cake, $NaHSO_4$, would be formed in making 500 tons of nitric acid by heating sodium nitrate with an excess of sulfuric acid?

23. How many pounds of silver nitrate, $AgNO_3$, could be obtained by the evaporation of the solution resulting from the treatment of 450 grams of silver, Ag, with nitric acid?

24. 63 grams of ammonia are oxidized by means of a catalyzer to nitric acid, which is absorbed by 200 cc. of water. What is the weight of the nitric acid solution formed?

25. How many pounds of calcium cyanamide, CaN_2C, may be produced by the reaction between 1000 pounds of calcium carbide, CaC_2, and nitrogen at a high temperature and under pressure, in an electric furnace?

26. 1 ton of calcium cyanamide, 40% CaN_2C, is treated in a closed vessel with steam under pressure, so that the nitrogen present is all converted into ammonia. What weight, in pounds, of ammonia is formed?

27. 356 cc. of ammonia is decomposed by passing the gas through a red hot tube. What volume of hydrogen will be produced?

28. The density of commercial nitric acid is 1.40 compared with water. What is the weight in grams of 60 cc. of this acid?

29. What is the volume in cubic centimeters of 100 grams of nitric acid, specific gravity 1.42?

30. Nitric acid may be completely decomposed by heating it to 256° C. What volume of nitrogen dioxide, NO_2, may be produced by thus decomposing 10 grams of pure nitric acid?

31. Concentrated nitric acid, when heated with a reducing agent, oxidizes the reducing agent. If 2.6 grams of charcoal, 98% C., is heated in contact with concentrated nitric acid, what volume of carbon dioxide, CO_2, will be produced?

The Oxides of Nitrogen

There are five recognized oxides of nitrogen, of which only three are commonly prepared. These are: nitrous oxide, N_2O; nitric oxide, NO; and nitrogen dioxide, NO_2.

At temperatures below 22° C., two molecules of nitrogen dioxide, NO_2, unite to form one molecule of nitrogen tetroxide, N_2O_4, which is yellow in color as contrasted with the red brown color of the compound when heated above 22° C.

The other oxides are nitrogen trioxide, N_2O_3, which is considered to be the anhydride of nitrous acid, and nitrogen pentoxide, N_2O_5, which is considered to be the anhydride of nitric acid.

1. How much water would be formed in the decomposition of 200 grams of ammonium nitrate, 80% NH_4NO_3, to form laughing gas, N_2O?

2. How much nitric oxide, NO, by volume could be made by the action of nitric acid upon 4.284 kilograms of copper, Cu?

3. How much oxygen by volume would be used up in changing 750 liters of nitric oxide, NO, to nitrogen dioxide, NO_2?

4. What weight of pure ammonium nitrate, NH_4NO_3, is required to produce 100 liters of laughing gas?

5. How many grams of nitric acid, 70% HNO_3, are needed to yield 1 liter of nitric oxide, NO, gas by its reaction with copper, Cu?

6. What volume of nitrogen dioxide, NO_2, is formed by bringing together 10 liters of air, 20% oxygen, and 10 liters of nitric oxide, NO?

7. How many grams of laughing gas, nitrous oxide or N_2O, can be obtained from 30 grams of ammonium nitrate, NH_4NO_3, through heating?

8. How many grams of ammonium nitrate, NH_4NO_3, would be required to produce, by heating, 500 liters of laughing gas, N_2O?

9. How many liters of nitric oxide, NO, would be formed by the action of an excess of nitric acid upon 50 grams of copper, Cu?

10. What weight of nitric acid, 68% HNO_3, would be used up in the production of 400 liters of nitric oxide, NO, by its action upon copper, Cu?

11. What weight of copper nitrate, $Cu(NO_3)_2$, would result during the formation of 250 liters of nitric oxide, NO, by the action of nitric acid upon copper, Cu?

12. What volume of air would be used up in changing 700 liters of nitric oxide, NO, to nitrogen dioxide, NO_2? Assume that air is 20% oxygen by volume.

13. 10 grams of copper, Cu, is completely reacted upon by an excess of nitric acid, and the nitric oxide, NO, formed is oxidized to nitrogen dioxide, NO_2. What volume of nitrogen dioxide would result?

14. A dentist's tank contains 50 cubic feet of nitrous oxide, N_2O. How many pounds of ammonium nitrate, NH_4NO_3, must be decomposed by heating to produce this amount of gas? One cubic foot of nitrous oxide weighs 0.00254 lb. at standard temperature and pressure.

15. 754 cc. of nitric oxide, NO, is mixed with just enough oxygen to convert it completely to nitrogen dioxide, NO_2. What is the volume of nitrogen dioxide produced?

16. 800 cc. of nitrogen dioxide, NO_2, is absorbed in 10 grams of water and all of the nitrous acid, HNO_2, formed is oxidized to nitric acid by exposure to the air. What is the per cent strength of the nitric acid formed?

CHAPTER XIII. SULFUR

This element is one of the most important of all, industrially. This is true to such an extent that the amount of its principal compound, sulfuric acid, which a country uses serves as an index to the amount of manufacturing that the country does. We may well ponder the statement, "The civilization of a country is determined by the amount of sulfuric acid that it uses."

1. What weight of iron sulfide, FeS, would be needed to furnish 20 liters of hydrogen sulfide, H_2S, by its reaction with an acid?

2. Assuming that one fifth of the air by volume is made up of oxygen, what volume of air would be needed to completely consume 150 cubic feet of hydrogen sulfide, H_2S? What volume of sulfur dioxide, SO_2, would be formed?

3. What weight of copper sulfate, $CuSO_4$, would be formed in the preparation of 22.4 liters of sulfur dioxide, SO_2, by heating copper with concentrated sulfuric acid?

4. How much pure sulfur, S, would be used up in manufacturing 3 tons of pure sulfuric acid?

5. 640 pounds of sulfur, S, is converted into sulfuric acid, 85% H_2SO_4. How much water is used up all together in the process?

6. How many grams of hydrogen sulfide, H_2S, would be produced in the reaction between dilute sulfuric acid and 20 grams of iron sulfide, FeS?

7. What weight of sulfur dioxide, SO_2, would result from the complete combustion of 1000 liters of hydrogen sulfide, H_2S?

8. What volume of sulfur dioxide, SO_2, is produced if 30 grams of pure sulfuric acid is destroyed completely by heating it with copper?

9. Twenty tons of crude sulfur is used to produce 50 tons of sulfuric acid, 95% H_2SO_4. What is the degree of purity of the crude sulfur, expressed in per cent?

10. What volume of air, 20.9% oxygen, is required to oxidize 1 kilogram of sulfurous acid, 10% H_2SO_3, solution to sulfuric acid by slow oxidation?

11. What weight of sulfur trioxide, SO_3, is formed through the oxidation of 10 liters of sulfur dioxide, SO_2, by the catalytic aid of sponge platinum?

12. The density of commercial concentrated sulfuric acid, 98% H_2SO_4, is 1.84. How many cubic centimeters of this acid are required to neutralize 40 grams of potassium hydroxide in dilute solution?

13. What weight of sulfuric acid, 67% H_2SO_4, can be formed from 75 tons of pure sulfur, S?

14. How much water by weight would be removed by evaporation while converting 300 tons of crude sulfuric acid, 67% H_2SO_4, to concentrated sulfuric acid, 98% H_2SO_4?

15. What weight of sulfuric acid, 70% H_2SO_3, could be manufactured from 600 tons of zinc blende ore, ZnS, if the ore is 30% sulfur, S?

16. What volume of hydrogen sulfide, H_2S, could be obtained from 0.5 kilo of ferrous sulfide, FeS, by the action of an acid upon it?

17. What weight of sulfurous acid, 30% H_2SO_3, could be prepared by dissolving in water the sulfur dioxide, SO_2, formed as the result of burning 5 kilos of iron pyrites, FeS_2?

18. What weight of sulfuric acid, H_2SO_4, would be formed by the oxidation of 30 tons of sulfurous acid, H_2SO_3?

19. How many tons of pure zinc sulfide, ZnS, would be required to furnish enough sulfur, S, for the manufacture of 1000 tons of concentrated sulfuric acid, 98% H_2SO_4?

20. 60 grams of copper is heated with an excess of concentrated sulfuric acid and the evolved sulfur dioxide, SO_2, is collected and measured. What volume would this gas occupy at room temperature, 20° C., and standard pressure?

21. What weight of water would be used up in converting the sulfur trioxide, SO_3, formed by the oxidation, in the presence of a catalyzer, of 1 ton of pure sulfur, S, to sulfuric acid, H_2SO_4?

22. 2 tons of pure sulfur, S, is oxidized to sulfur dioxide, SO_2, then to sulfur trioxide, SO_3, and is then made into commercial sulfuric acid, 67% H_2SO_4. What is the total weight of water used up in the process?

23. What weight of sodium sulfite, Na_2SO_3, would be required to form 30 liters of sulfur dioxide, SO_2, by heating it with sulfuric acid?

24. What volume of hydrogen sulfide, H_2S, at a temperature of 30° C. and a pressure of 700 mm. of mercury, would be formed if 50 grams of ferrous sulfide, FeS, were completely reacted upon by hydrochloric acid?

25. At standard temperature and pressure 4.37 volumes of hydrogen sulfide, H_2S, will dissolve in one volume of water. What weight of saturated solution of hydrogen sulfide would result by dissolving the gas given off in the complete reaction between 25 grams of ferrous sulfide, FeS, and an excess of sulfuric acid?

26. What volume of sulfur dioxide, SO_2, would result from the complete combustion of 500 liters of hydrogen sulfide, H_2S?

27. What weight of copper would be required to produce, by its action upon hot concentrated sulfuric acid, 90 liters of sulfur dioxide, SO_2?

28. 70 grams of copper is completely reacted upon by hot concentrated sulfuric acid and the resulting solution evaporated to dryness. What weight in grams of crystallized copper sulfate, $CuSO_4 \cdot 5\,H_2O$, could be obtained from the product?

29. How many tons of sulfuric acid, 98% H_2SO_4, could be manufactured from 3000 tons of an ore which contains 30% of sulfur, S, in combination?

30. Hydrogen sulfide gas, H_2S, is passed into an alkaline solution of cadmium chloride, $CdCl_2$. The precipitate of cadmium sulfide, CdS, is dried and found to have a weight of 2.3 grams. What volume of hydrogen sulfide entered into the reaction?

31. How many grams of pure sulfuric acid may be obtained by the proper oxidation of the sulfur, S, contained in 100 liters of pure hydrogen sulfide, H_2S?

32. 3.2 grams of sulfur is mixed with 7.2 grams of iron, Fe, and the mixture heated until the reaction is complete. How much ferrous sulfide, FeS, is formed? Which element will be only partially used up and how much of it will remain uncombined?

33. Hydrogen sulfide has a solubility at ordinary temperature and pressure of 3 volumes in 1 of water. How much ferrous sulfide, FeS, will be required to produce enough hydrogen sulfide to saturate 3 liters of water?

34. Sulfur dioxide, SO_2, has a solubility at ordinary room temperature of 60 volumes in 1 volume of water. How many grams of sulfur must be burned in air to produce the sulfur dioxide necessary to saturate 100 cc. of water?

35. How many grams of sulfuric acid, 98% H_2SO_4, must be decomposed, by heating it with copper, to yield enough sulfur dioxide to saturate 1 liter of water? Under ordinary conditions 60 volumes of sulfur dioxide will dissolve in 1 volume of water.

36. How much sodium sulfite, Na_2SO_3, must be acted upon by dilute sulfuric acid in order to produce 10 grams of pure sulfurous acid, H_2SO_3?

37. How much sodium sulfite, Na_2SO_3, will be produced by passing 10 liters of sulfur dioxide into a solution of sodium hydroxide?

38. What weight of water must be added to 30 grams of pure sulfur trioxide, SO_3, to convert all of it into sulfuric acid, 95% H_2SO_4?

39. How much sulfuric acid, 96% H_2SO_4, is required to produce 10 grams of crystallized sodium sulfate, $Na_2SO_4 \cdot 10\,H_2O$, by its reaction with sodium carbonate, Na_2CO_3, in solution?

40. How much crystallized potassium sulfate, K_2SO_4, can be prepared by the reaction between 30 cc. of sulfuric acid, 98% H_2SO_4, density 1.84, and potassium chloride, KCl, at a red heat?

41. What weight, in pounds, of sodium hydrogen sulfate, $NaHSO_4$, is needed to produce the same acid effect as 10 pounds of sulfuric acid, 95% H_2SO_4?

CHAPTER XIV. THE CHLORINE FAMILY

Chlorine — Bromine — Iodine

The elements of this family may be prepared from their hydrogen compounds by the general method of oxidation. The source of the oxygen is immaterial so long as that element is present in the nascent or free atomic condition. The element *fluorine*, however, can only be prepared by electrolysis of the fused salt, since its affinity for hydrogen is too strong for oxygen to break. Fluorine is not strictly a member of this family although it does belong to the same group.

1. How much chlorine, Cl_2, by weight could be produced by the oxidation of 73 grams of hydrogen chloride, HCl?

2. 1 ton of salt, 90% NaCl, is treated with an excess of sulfuric acid and the hydrogen chloride gas, HCl, is absorbed in water. How many pounds of commercial hydrochloric acid, 37% HCl, can be made?

3. How many grams of manganese dioxide, MnO_2, would be needed, with sulfuric acid, to use up 206 grams of sodium bromide, NaBr, in the usual laboratory method of making bromine, Br_2?

4. How much sulfuric acid by weight would be required to produce 254 grams of iodine, I_2, from an iodide with the aid of an oxidizing agent?

5. How much bromine water, 3% Br_2, can be prepared from 1 gram of potassium bromide, KBr, by the action upon it of sulfuric acid and an oxidizing agent?

6. What weight of iodine, I_2, is displaced in potassium iodide, KI, solution by the action of 10 grams of chlorine water, 5% Cl_2?

7. What weight of chlorine, Cl_2, could be formed by the oxidation of 50 grams of hydrochloric acid, 38% HCl?

8. What volume of chlorine, Cl_2, measured at 30° C., would be formed through heating 1 kilo of hydrochloric acid, 37% HCl, with manganese dixiode, MnO_2?

9. What volume of chlorine, Cl_2, would result from the electrolysis of 3 kilos of fused sodium chloride, NaCl?

Gas Problems Involving Reacting Volumes

Experiments upon gases have shown that the molecules of one gas at a given temperature and pressure take up as much space as the molecules of another gas under the same conditions of temperature and pressure. If we consider oxygen we find that 32 grams of this gas, measured at 0° C. and 760 mm. of mercury pressure, occupy 22.4 liters. Likewise we find that 70.92 grams of chlorine will, under the same conditions, occupy 22.4 liters. Also, 64.07 grams of sulfur dioxide, 36.46 grams of hydrogen chloride, or 20 grams of hydrogen fluoride, would be found to occupy 22.4 liters if measured under these same conditions of temperature and pressure.

In other words, 22.4 liters of any gas, if measured at standard temperature and pressure, would be found to weigh the number of grams expressed in its molecular weight, or, since grammolecular weights of gases indicate equal numbers of molecules, then 22.4 liters of one gas contain as many molecules as 22.4 liters of another gas.

From this we can readily see that one liter of a gas must contain as many molecules as one liter of any other gas under the same conditions of temperature and pressure. This leads us up to the law of gases as expressed by Avogadro in his hypothesis, in which he states that equal volumes of any two gases contain the same number of molecules, if existing conditions are the same.

Now, in a reaction between gases we can see that if one liter of hydrogen combines with one liter of chlorine then one molecule of hydrogen must combine with one molecule of chlorine. In the same way, if one molecule of nitrogen combines with three molecules of hydrogen, then one liter of nitrogen will combine with three liters of hydrogen. This fact greatly simplifies the solving of problems based upon the reactions between gases where the volumes are known, since it is not necessary to deal with the weight relationship at all.

The following problem may serve as an illustration:

What volume of oxygen would be required for the complete combustion of 500 liters of hydrogen sulfide? The reaction in this case is expressed by the equation:

$$2\,H_2S + 3\,O_2 \longrightarrow 2\,H_2O + 2\,SO_2.$$

The ratio of the hydrogen sulfide molecules to the oxygen molecules is 2 : 3. The combining volumes of these same gases must also be in the ratio 2 : 3. The proportion therefore becomes: $2 : 3 :: 500 : x$. $2\,x = 1500$ and $x = 750$.

The required answer is therefore *750 liters* of oxygen.

1. By means of an electric spark, 500 liters of hydrogen was exploded with chlorine to form hydrogen chloride, HCl. What volume of chlorine took part in the reaction and what was the volume of hydrogen chloride that was formed?

2. What weight of hydrogen chloride, HCl, would be formed by heating 300 pounds of sodium chloride, NaCl, with sulfuric acid?

3. 80 grams of sodium chloride, $NaCl$, is heated with an excess of sulfuric acid. What volume of hydrogen chloride, HCl, would result?

4. What weight of hydrochloric acid, 38% HCl, could be manufactured from 10 tons of sodium chloride, $NaCl$, by heating it with an excess of sulfuric acid?

5. What volume of fluorine, F_2, would be formed by the electrolysis of 300 grams of liquid hydrogen fluoride, HF?

6. What weight of hydrofluoric acid, 50% HF, could be prepared from 200 grams of calcium fluoride, CaF_2, by heating it with concentrated sulfuric acid in a lead or platinum dish?

7. What weight of bromine, Br_2, would be formed from 5 grams of potassium bromide, KBr, by heating it with manganese dioxide, MnO_2, and sulfuric acid?

8. What weight of iodine, I_2, could be released in a solution of potassium iodide, KI, by passing 900 cc. of chlorine, Cl_2, through the solution?

9. Calculate the percentage composition of hydrogen fluoride, by volume.

Calculate the percentage composition of hydrogen fluoride, by weight.

10. What weight of calcium fluoride, CaF_2, would be required in the production of 80 liters of hydrogen fluoride, HF, by the action of concentrated sulfuric acid upon it?

11. What volume of fluorine, F_2, would result from the electrolysis of 200 grams of hydrogen fluoride, HF?

12. What weight of water would be formed by the action of hydrogen fluoride, HF, upon 40 grams of silica, SiO_2?

13. What volume of chlorine, Cl_2, could be formed from 5 kilograms of hydrochloric acid, 37% HCl, by the action of an oxidizing agent upon it?

14. Metallic sodium, Na, is burned in chlorine, Cl_2, until the chlorine is completely reacted upon. The volume of the chlorine was 60 liters, measured at a temperature of 25° C. What weight of sodium chloride would result from the combustion?

15. A current of electricity is passed through a solution of sodium chloride, NaCl, until 22.4 liters of chlorine, Cl_2, is produced. What weight of sodium hydroxide would be left in the solution?

16. How many grams of manganese dioxide, MnO_2, would be needed to completely oxidize 600 grams of hydrochloric acid, 20% HCl, in the preparation of chlorine, Cl_2?

17. What weight of hydrochloric acid, 25% HCl, would be formed by dissolving in water the gas obtained by exploding a mixture of 400 liters of chlorine, Cl_2, and 500 liters of hydrogen, H_2?

18. How many tons of hydrochloric acid, 35% HCl, could be manufactured from 800 tons of sodium chloride, NaCl, by heating the salt with an excess of sulfuric acid?

19. What weight of anhydrous sodium sulfate, Na_2SO_4, would be formed as a by-product in the manufacture of 4000 tons of hydrochloric acid, 30% HCl, by the reaction between sodium chloride and sulfuric acid, at a high temperature?

20. What weight of bromine, Br_2, could be made by heating 60 grams of sodium bromide, NaBr, with manganese dioxide, MnO_2, and sulfuric acid?

21. How many liters of hydrogen bromide, HBr, would result from the reaction between 150 grams of sodium bromide, NaBr, and dilute sulfuric acid, without taking into consideration the fact that part of the product is oxidized by the sulfuric acid?

22. Pure hydrogen bromide, HBr, is best formed by moistening red phosphorus with water and allowing bromine, Br_2, to fall, drop by drop, into the paste. The reaction is as follows:

$$2\,P + 3\,Br_2 + 6\,H_2O \longrightarrow 2\,H_3PO_3 + 6\,HBr.$$

What volume of hydrogen bromide could be formed in this way from 200 grams of bromine?

23. An important source of iodine is potassium iodate, KIO_3, which forms approximately 0.2% of crude Chile saltpeter. What weight of the element is present in one ton of the ore?

24. What weight of iodine, I_2, could be obtained from 500 pounds of sodium iodide, NaI, by the English method of heating a mixture of sodium iodide, manganese dioxide, MnO_2, and sulfuric acid?

25. What weight of chlorine, Cl_2, would be required to produce 1 kilogram of iodine, I_2, by the French method of displacing the iodine in a solution of sodium iodide, NaI, with chlorine?

26. A solution of hydrogen iodide, hydriodic acid, HI, can be prepared by passing hydrogen sulfide, H_2S, into water in which iodine is suspended. The sulfur is precipitated while the gas formed is dissolved in the water. What weight of iodine, I_2, would be required to prepare 3000 grams of hydriodic acid, 25 per cent HI?

27. What is the per cent of chlorine in sodium chloride, NaCl?

28. What volume of oxygen will be released from water by the action of 1 gram of chlorine, with the aid of sunlight?

29. What volume of hydrogen fluoride, HF, will be formed by the reaction between concentrated sulfuric acid and 1 gram of calcium fluoride, CaF_2?

30. How many cubic centimeters of hydrogen will combine exactly with 166 cc. of chlorine when a mixture of the gases is exposed to direct sunlight? What volume of hydrogen chloride, HCl, will be produced?

31. Under ordinary conditions 430 volumes of hydrogen chloride dissolve in one volume of water. How much salt, NaCl, must be heated in solution with sulfuric acid to produce enough hydrogen chloride, HCl, to saturate 200 cc. of water under ordinary conditions?

32. How many pounds of potassium chlorate, $KClO_3$, may be produced from 100 pounds of potassium chloride, KCl, by its reaction with calcium chlorate, $Ca(ClO_3)_2$?

33. How many cubic centimeters of a normal solution of potassium bromide, KBr, are required to precipitate 0.12 gram of silver bromide, AgBr, from a solution of silver nitrate, $AgNO_3$?

34. Chlorine, Cl_2, may be prepared, according to Deacon's process, by passing a mixture of air and hydrogen chloride, HCl, over heated copper chloride, $CuCl_2$, mixed with asbestos. How many liters of chlorine could be obtained in this way from 60 liters of hydrogen chloride?

35. How many liters of air, 20% oxygen, are required, by Deacon's process, to oxidize 50 liters of hydrogen chloride, HCl, to chlorine, Cl_2, and water?

36. What weight of sodium chloride, NaCl, is required to furnish sufficient chlorine, Cl_2, to change 30 grams of calcium hydroxide, $Ca(OH)_2$, completely into bleaching powder, $CaOCl_2$?

37. The amount of chlorine present in a compound is often determined by precipitating and weighing it in the form of silver chloride, AgCl. What is the per cent strength of a solution of sodium chloride, NaCl, if 10 g. of the solution produces 0.26 gram of silver chloride, AgCl, by its reaction with silver nitrate, $AgNO_3$, in solution?

CHAPTER XV. THE DETERMINATION OF MOLECULAR AND ATOMIC WEIGHTS

Molecular Weights

In order to determine the formula of any chemical compound, it is necessary to know its molecular weight. Furthermore, the same knowledge is required to determine the atomic weights of the elements.

Gases lend themselves most readily to molecular weight determination because of a principle established by Avogadro, an Italian physicist. This principle may be stated thus: Equal volumes of gases, measured under the same conditions of temperature and pressure, contain equal numbers of molecules. This holds good, regardless of the composition of the gas.

In accordance with this principle, we perceive that one liter of nitrogen contains the same number of molecules as does one liter of oxygen measured under the same conditions. Therefore the weight of one molecule of nitrogen bears the same ratio to that of one molecule of oxygen as does the weight of one liter of nitrogen to that of one liter of oxygen, measured under the same conditions.

We find by actual experiment that one liter of nitrogen weighs 1.252 grams, while one liter of oxygen weighs 1.429 grams. The ratio of weight between one liter of nitrogen and one liter of oxygen is then $\frac{1.252}{1.429}$. The ratio between the molecular weights is the same as between the weights of equal volumes so that the same ratio, $\frac{1.252}{1.429}$, represents the relative weights of the nitrogen and oxygen molecules.

It has been determined that one molecule of oxygen contains two atoms of oxygen. The atomic weight of oxygen is 16. Therefore the molecule of oxygen weighs 2×16 or 32. According to the ratio just determined, the molecule of nitrogen weighs $\frac{1.252}{1.429} \times 32$ or 28.03. This, then, is the molecular weight as determined for nitrogen.

If we take as many grams of a compound as there are atomic units in the molecular weight, we have the *gram-molecular weight* of the compound. One gram-molecular weight of oxygen is therefore 32 grams. One gram-molecular weight of any substance always represents the same number of molecules, regardless of the kind of substance. Equal numbers of molecules of gases, measured under the same conditions, indicate equal volumes of the gases.

It will be seen, then, that gram-molecular weights of gases always have the same volume, when measured under the same conditions, regardless of the kind of gas. Usually standard conditions of temperature and pressure are chosen for comparisons of gas volumes, as a matter of convenience. Under such conditions, one gram-molecular weight of oxygen, 32 grams, occupies 22.4 liters. Or, we may state it thus: 22.4 liters of oxygen represents one gram-molecular weight of oxygen.

Now, since gram-molecular weights of gases always represent the same volume, under standard conditions, we may determine the molecular weight of any substance in the form of a gas, by finding the weight in grams of 22.4 liters of the gas.

Example: 1 liter of nitrous oxide, N_2O, under S. T. P. weighs 1.966 grams. What is its molecular weight?

Method: Determine the weight of 22.4 liters of the gas.

$$22.4 \times 1.966 = 44.04.$$

The molecular weight of nitrous oxide, as calculated, is therefore *44.04*.

Example: 250 cc. of a gas measured at 17° C. and 740 mm. of mercury pressure has a weight of 0.32 gram. Determine the molecular weight of the gas.

Method: First, reduce the volume to S. T. P.

$$\overset{25}{\cancel{250}}\times\frac{\overset{37}{\cancel{740}}}{\underset{38}{\cancel{760}}}\times\frac{273}{\underset{29}{\cancel{290}}}=229.1 \text{ cc. measured under S. T. P.}$$

Since there are 1000 cc. in one liter, $\frac{1000}{229.1} \times 0.32 \text{ g.} = 1.3967 \text{ g.}$, the weight of one liter of the gas, and

22.4×1.3967 g. $= 31.29$ g., the gram-molecular weight of the gas. The molecular weight is then *31.29*.

TO DETERMINE THE WEIGHT OF ONE LITER OF A GAS FROM THE FORMULA

The factor 22.4 may be used to determine the weight of one liter of any gas, under S. T. P., provided the formula of the molecule of the gas is known. We have just seen that 22.4 times the weight in grams of one liter of a gas is its gram-molecular weight. Therefore, the gram-molecular weight of any gas divided by the factor 22.4 gives us the weight of one liter of the gas under S. T. P.

Example: What is the weight of one liter of carbon dioxide, CO_2, at S. T. P.?

Method: The molecular weight of carbon dioxide is the sum of 12 (the atomic weight of carbon) plus 32 (twice the atomic weight of oxygen) which is then 44.

$\frac{44}{22.4}=1.96$ grams, weight of one liter of carbon dioxide.

The molecular weights of substances in the liquid or in the solid state may be determined by several methods which will not be discussed here. Any standard college textbook of Chemistry will indicate some of the methods.

Atomic Weights

To determine the weight of an atom, it is necessary to know: (1) the exact percentage composition of several compounds of the element; (2) the molecular weight of each of these compounds; (3) the equivalent weight of the element, accurately determined. We then multiply the molecular weight of each compound by the per cent of the particular element in that molecule, the number obtained being the total weight, expressed in atomic units, of the element in the molecule.

The numbers found in this way will be even multiples or parts of each other. Since no weight so determined can be less than the weight of one atom of the element in question, the smallest number found in investigating the composition of a large number of compounds of the element is probably the weight of a single atom of the element. This number is taken as the approximate atomic weight of the element.

To determine the exact atomic weight, it is necessary to know the equivalent weight of the element. The equivalent weight of an element is the weight of an element which combines exactly with a given weight of another element taken as standard. The most convenient standard is oxygen, and the equivalent of an element may be defined as the weight of the element which combines exactly with one gram-atomic weight of oxygen, 16 grams.

The multiple of the equivalent weight of an element which most nearly approaches the smallest approximate atomic weight, as previously determined, is taken as the true atomic weight of the element.

Example: It is desired to determine the atomic weight of sulfur from the data as follows:

Compound	Molecular Weight	Per Cent Sulfur by Analysis	Atomic Units of Weight Due to Sulfur
A	64.	50.	32.
B	80.	40.	32.
C	34.	94.1	31.99
D	72.	44.4	31.97
E	95.6	33.5	32.03
F	97.4	32.9	32.04
G	336.	28.6	96.1

In the third column are shown the products of the molecular weights of the compounds by the per cent of sulfur in those compounds, the product in each case representing the number of atomic units of weight of sulfur present in the compound. By inspection of this table it will be seen that the amounts of sulfur present are either very nearly 32 units or an even multiple of 32 units (*i.e.* 96 units). The atomic weight must then be nearly 32, since this is the smallest value found for sulfur.

Now, in the compound sulfur dioxide, it is found that exactly 16.03 grams of sulfur are united with 16 grams of oxygen. 16.03 is then the equivalent weight of sulfur. 2×16.03 gives 32.06, the multiple of the equivalent most nearly approaching the value of 32, previously determined. The exact atomic weight of sulfur is thus established as 32.06.

1. One liter of a gas weighs 3.616 grams. Find its molecular weight.

2. 100 cc. of a gas, at 30° C. and under a pressure of 760 mm. of mercury, weighs 0.1289 gram. Calculate its molecular weight.

3. A certain gas is composed of 15.79% carbon and 84.21% sulfur. One liter of it weighs 3.393 grams. Find the molecular weight of the gas and its formula.

4. A gas has a specific gravity, referred to air, of 1.806. Calculate its molecular weight.

5. 1 liter of a gas weighs 1.162 grams. Find its molecular weight.

6. What is the weight of 1 liter of propane gas, C_3H_8, measured at a temperature of 20° C. and a pressure of 740 mm. of mercury?

7. 253 cc. of a certain gas mixture weighs 0.32637 gram. What is the average molecular weight of the gas mixture?

8. The combining weight of an element, *X*, is found to be 6. By analysis it is found present in several gaseous compounds as follows, a molecule of each compound containing 1 atom of *X*. Find the atomic weight of the element.

Compound	Weight of 1 Liter of Gas	Per Cent *X* in Molecule
A	0.7168	75.
B	1.9768	27.3
C	1.2504	42.8
D	1.1607	46.16

9. 1 liter of a certain gas weighs 2.9266 grams. From this calculate its molecular weight.

10. 1 liter of a gas weighs 1.2504 grams. What is its molecular weight?

11. 560 cc. of a gas weighs 0.431648 gram. Find the molecular weight.

12. 250 cc. of a gas weighs 0.1792 gram. Calculate its molecular weight.

13. 0.9 liter of a certain gas weighs 1.12563 grams. Find the molecular weight.

14. 2100 cc. of a gas weighs 3.44358 grams. What is its molecular weight?

15. 300 cc. of air weighs 0.38784 gram. From this calculate the average molecular weight of air.

16. The specific gravity, with respect to air, of a certain gas is 0.6907. What would 1 liter of the gas weigh?

17. The specific gravity of a gas is 1.53. Find the molecular weight of the gas.

18. Find the specific gravity of a gas which has a molecular weight of 34.

19. The molecular weight of a certain gas is 28. Calculate its specific gravity.

20. Calculate the vapor density, with respect to hydrogen, of a certain gas if 230 cc. of the gas weighs 0.766 gram.

21. 0.5772 gram of a certain gas, at a temperature of 20° C. and a pressure of 740 mm. of mercury, measures 220.45 cc. Calculate its molecular weight.

22. 0.23025 gram of a gas measures, at a temperature of −20° C. and pressure of 800 mm. of mercury, 132.06 cc. Find its molecular weight.

23. The molecular weight of a gas is 46. Find the weight of 1 liter.

24. The molecular weight of a gas is 70.92. What does 1 liter of it weigh?

25. The molecular weight of a gas is 20.008. What would 1 liter of it weigh if it were measured under a pressure of 750 mm. of mercury and at a temperature of 25° C.?

26. 1 kiloliter of a gas weighs 1440 grams. From this calculate its molecular weight.

27. A certain gas has the following composition: Oxygen, 72.73%; carbon, 27.27%. What is its simplest formula?

28. An analysis of 100 grams of a certain gas yielded the following results: Carbon, 64.87%; oxygen, 21.62%; hydrogen, 13.51%. Find its simplest formula.

29. A substance upon analysis was found to consist of: Nitrogen, 8.28%; silver, 63.31%; oxygen, 28.41%. Find its simplest formula.

30. A substance, when analyzed, was found to have the following composition: Potassium, 44.87%; oxygen, 36.73%; sulfur, 18.4%. Find its simplest formula.

31. The molecular weight of a certain gas is 26. It is composed of: Carbon, 92.25%; hydrogen, 7.75%. What is its formula?

32. A substance has a molecular weight of 60.4. It is composed of: Oxygen, 52.98%; silicon, 47.02%. Find its formula.

33. 250 cc. of a certain gas weighs 0.91125 gram. Its composition is: Bromine, 98.75%; hydrogen, 1.25%. From this calculate its formula.

34. A gas, upon analysis, is found to consist of: Oxygen, 69.55%; nitrogen, 30.45%. 360 cc. of the gas weighs 1.4922 grams. What is its formula?

35. A gas is found to have a specific gravity of 2.5. What is its molecular weight?

36. The gas, hydrogen telluride, H_2Te, has a molecular weight of 129.6. What is the weight of 1 liter of the gas?

37. 250 cc. of the gas, hydrogen selenide, H_2Se, weighs 0.90 gram under standard conditions of temperature and pressure. What is its molecular weight?

38. What volume will be occupied, at a pressure of 760 mm. of mercury and a temperature of 20° C., by 2 grams of a gas which has a molecular weight of 44.02?

39. Carbon dioxide has the formula, CO_2. What volume would 2.5 grams of this gas occupy if it were collected over water, at a temperature of 25° C. and a pressure of 750 mm. of mercury? Neglect the decrease in volume that would result from partial solution of the gas under these conditions.

40. How many grams weight are there in 640 cc. of a gas which has a molecular weight of 44?

41. 0.5 gram of hydrogen is collected over water. What volume does it occupy at standard temperature and pressure?

42. 3.6 grams of oxygen is collected over water, at a temperature of 18° C. and a pressure of 740 mm. of mercury. What volume does the oxygen occupy?

43. The total pressure exerted by the gas, in a mixture of gases, is due to the sum of the pressures of the individual gases present. In a mixture of oxygen, 20%, and nitrogen, 80%, measured at a pressure of 760 mm. of mercury, what pressure is due to the oxygen? To the nitrogen?

44. 560 cc. of nitrogen, collected over water, is measured under a pressure of 750 mm. of mercury and a temperature of 20° C. What pressure is exerted by the nitrogen alone?

45. A vessel containing oxygen, collected over water, is found to contain 1 liter of gas, at a temperature of 16° C. and standard pressure. What is the weight of oxygen present at this temperature?

46. 360 cc. of a gas was collected over water, at a temperature of 18° C. and under a pressure of 740 mm. of mercury. Reduce the volume to standard conditions of temperature and pressure.

47. A sample of natural gas is found, by analysis, to have the following per cent composition by volume: Methane, 90%; carbon dioxide, 1%; and nitrogen, 9%. What is the pressure exerted by the carbon dioxide in the mixture when the gas is under standard conditions of temperature and pressure?

48. 400 liters of a gas mixture is collected over water at 20° C. and 750 mm. of mercury pressure. The mixture is then run through a drying agent and measured at standard temperature and pressure when, by analysis, it is found to have the following composition by volume: Oxygen, 50%; nitrogen, 25%; and hydrogen, 25%. What would be the volume of each of these gases under the new conditions?

CHAPTER XVI. CARBON AND SIMPLE COMPOUNDS OF CARBON

Carbon

The element carbon is the main constituent of practically all fuels. On account of its easy availability and its high affinity for oxygen it is the most widely used reducing agent. This element forms many very complex compounds which will be mentioned later.

1. How many tons of carbon dioxide, CO_2, would be driven into the air by the complete combustion of 1 ton of Pennsylvania bituminous coal if it is 75% carbon?

2. How many cubic feet of air would be required to completely consume 500 cubic feet of water gas; assuming that it is composed of carbon monoxide, CO, and hydrogen, H_2, 50% of each by volume? Take air as containing 20% oxygen by volume.

3. How much steam by weight could be decomposed by the carbon, C, in 1 ton of coke if the carbon makes up 86.3% of the coke by weight?

4. What weight of calcium carbide, CaC_2, can be made from 1 kilogram of pure calcium oxide, CaO, by heating it with coke, C, in an electric furnace?

5. How many liters of acetylene, C_2H_2, can be produced by the action of water on 64 grams of calcium carbide, CaC_2?

6. What volume of pure oxygen would be needed for the complete combustion of 25 cubic feet of acetylene, C_2H_2, in an oxy-acetylene torch?

7. A 1-gram sample of coke is burned with oxygen and the carbon dioxide, CO_2, produced is absorbed in a tube containing concentrated potassium hydroxide solution. During the process the absorption tube gains 3.22 grams in weight. What is the per cent of carbon, C, in the sample of coke?

8. A sample of pure water gas is analyzed and found to contain 6.0 grams of carbon monoxide, CO. What is the volume of the sample?

9. What volume of carbon monoxide, CO, may be produced by passing 1 liter of air over red-hot carbon in a combustion tube? Regard the air as being 20% oxygen by volume for purposes of easy calculation.

10. How many grams of carbon are required to produce 60 grams of calcium carbide, CaC_2, by fusion with calcium oxide, CaO, in an electric furnace?

11. How many pounds of pure sulfur are required to produce 800 pounds of carbon disulfide, CS_2, in the electric furnace, if all of the sulfur is utilized?

12. How many pounds of carbon dioxide are produced in the burning of 1000 cubic feet of a natural gas which has the volumetric composition: Methane, CH_4, 90%; nitrogen, 10%?

13. What weight of carbon would be required to completely reduce 1000 pounds of copper oxide, CuO? Assume that the gaseous product formed is all CO_2.

14. What weight of calcium carbonate, $CaCO_3$, would be formed through collecting in limewater, $Ca(OH)_2$, the carbon dioxide which would result from the reduction of 75 pounds of cuprous oxide, Cu_2O?

15. What volume of air would be used up in the complete combustion of 800 liters of carbon monoxide, CO? Take air as one fifth oxygen by volume.

16. What weights of sulfur and carbon would be needed to yield 5000 pounds of carbon disulfide, CS_2, by their reaction in an electric furnace? Assume that there is a 3% waste in the process.

17. What volumes of carbon dioxide and sulfur dioxide would result from the complete combustion of 15 kilograms of carbon disulfide, CS_2?

18. At 15° C. and a pressure of almost ten atmospheres, pure water will dissolve fifteen times its own volume of carbon dioxide, forming soda water. How many kilograms of soda water of this strength could be made by dissolving in water the carbon dioxide obtained from 5 kilograms of marble, $CaCO_3$, if it is completely reacted upon by hydrochloric acid?

19. What volume of carbon dioxide would be given off in the raising of a mass of dough containing 160 grams of a baking powder of the following composition: Cream of tartar or potassium acid tartrate, 50%; dried corn starch, 25%; baking soda or sodium bicarbonate, 25%? What volume would this gas occupy at a "rising" temperature of 80° F.?

20. What volume of air would be needed for the complete combustion to carbon dioxide of one ton of coke, 99% C.? Take air as one fifth oxygen.

21. The poisonous gas cyanogen, C_2N_2, may be formed by the union of carbon and nitrogen at high temperatures. What weight of pure carbon would be needed to form 300 liters of the gas in this way?

22. What weight of coke, 90% C., would be required to yield 1000 cubic feet of water gas, by its reaction, when incandescent, with steam?

23. In the complete combustion to carbon dioxide and water of 800 cubic meters of water gas, what volume of air would be required to furnish the necessary oxygen if it is 20% oxygen by volume?

24. A sample of graphite weighing 3 grams is burned in oxygen, and the carbon dioxide formed is absorbed in a tube containing potassium hydroxide, KOH. During the absorption the potassium hydroxide gains 10.9 grams in weight. What is the per cent of carbon in the graphite sample?

25. A fire extinguisher is charged with 200 grams of sodium bicarbonate, $NaHCO_3$, and sufficient sulfuric acid to react with it completely. How many liters of carbon dioxide, at ordinary room temperature of 20° C., would be produced by their reaction?

26. Find the percentage composition of methane, CH_4, by weight.

27. How many cubic feet of carbon dioxide are produced by the complete combustion of 1,000,000 cubic feet of Ohio natural gas, the composition of which is approximately: Hydrogen, 0.9%; methane, or CH_4, 89.5%; the unsaturated hydrocarbon, C_6H_6, 9.3%; carbon dioxide, nitrogen, oxygen, and other non-combustibles, 0.3%?

28. How many pounds of carbon dioxide would result from the complete combustion of 1000 cubic feet of methane, CH_4?

29. What volume of air would be necessary to furnish the carbon dioxide that would be needed to "set" the calcium hydroxide, $Ca(OH)_2$, in 1 ton of mortar which is 20% calcium hydroxide by weight?

30. What weight of carbon disulfide, CS_2, could be made from 960 pounds of sulfur, by passing sulfur vapor over white-hot coke in an electric furnace?

The Hydrocarbons

1. How much methane, CH_4, by volume could be completely consumed by the oxygen in 200 cubic feet of air if the oxygen makes up one fifth of the air by volume?

2. Calculate the volumes at S. T. P. of the products that would be formed by the complete combustion of 76 grams of methyl alcohol, CH_3OH.

3. What volume of oxygen would be required to burn to carbon dioxide and water 46 grams of ethyl alcohol, C_2H_5OH?

4. 10 cc. of ether, $(C_2H_5)_2O$, is burned completely and the carbon dioxide formed is absorbed in potassium hydroxide solution in a weighing tube. What is the increase in weight of the weighing tube? Consider ether as having a specific gravity of 0.719.

5. What weight of calcium hydroxide, $Ca(OH)_2$, will result during the preparation of 20 liters of acetylene, C_2H_2, by the action of calcium carbide, CaC_2, on water?

6. A bicycle lamp burning acetylene consumes 8000 cc. per hour. What weight of calcium carbide, 80% CaC_2, is required to produce enough gas by its action with water to keep the lamp burning six hours?

7. 100 liters of ethane, C_2H_6, are burned and the water formed is caught in an absorption tube containing calcium chloride, $CaCl_2$. What is the increase in weight of the absorption tube?

8. How many liters of oxygen would be consumed in the complete combustion of 800 liters of methane, CH_4?

9. In the preparation of 500 liters of acetylene, C_2H_2, by the action of water upon calcium carbide, CaC_2, what weights of water and calcium carbide would enter into the reaction?

10. What weight of water could be heated from 0° C. to 100° C. by the combustion of the acetylene, C_2H_2, obtained from one pound of calcium carbide, CaC_2? 1 gram molecular weight of acetylene yields 303,660 calories of heat in burning.

11. In the oxy-acetylene blowpipe when used for cutting, the mixture used is 1 volume of acetylene to 1.7 volumes of oxygen. How many kilograms of water could be heated from the freezing to the boiling point by the heat given off by its flame in one minute if it cuts through 12 inches of metal per minute and uses 0.033 cu. ft. of C_2H_2 per inch of cut?

12. 1 gram atomic weight of carbon gives while burning 96,480 calories of heat. 1 gram atomic weight of hydrogen gives out while burning 28,800 calories of heat. 1 gram molecular weight of acetylene yields, while decomposing under the influence of heat, 53,100 calories of heat. How much water could be heated from the freezing to the boiling point by the combustion of 1 cubic meter of acetylene?

13. How many kilograms of water could be heated from the freezing to the boiling point by the combustion of 1 cubic foot of acetylene?

CHAPTER XVII. THE PHOSPHORUS FAMILY

The members of this family are closely related to nitrogen, and all of them, with the possible exception of bismuth, form hydrides with formulas similar to that of ammonia. The members of this family usually show acid properties in combination. They all form important alloys with the metals.

1. What weight of pure phosphorus could be manufactured from 600 pounds of calcium phosphate, $Ca_3(PO_4)_2$, by heating it with sand and carbon in an electric furnace?

2. An ore contains 62.38% of arsenopyrite, FeAsS. What weight of arsenic could be obtained by heating 261.3 pounds of this ore, in the absence of air?

3. What weight of an ore containing 23.46% of stibnite, Sb_2S_3, would be needed to furnish 240 pounds of pure antimony, Sb, by fusion with iron?

4. How many grams of charcoal would be required to reduce 198 grams of arsenious oxide, As_2O_3? How many liters of carbon dioxide would be formed in the reaction?

5. On the assumption that just 90% of the phosphorus is recovered, what weight of phosphorus may be obtained from 150 grams of calcium phosphate, $Ca_3(PO_4)_2$, by heating it with sand and carbon in an electric furnace?

6. What weight of arsenious oxide, As_2O_3, would be formed in the combustion of 1 gram of arsenic, 95% As?

7. What weight of bismuth hydroxide, $Bi(OH)_3$, would be produced by the reaction between ammonium hydroxide and 0.15 gram of bismuth chloride, $BiCl_3$, in solution?

8. What volume of arsine, AsH_3, may be prepared by the action of nascent hydrogen upon 0.3 gram of arsenious oxide, As_2O_3, in acid solution?

9. What weight of pure carbon would be required to produce 300 pounds of phosphorus by its reaction with calcium phosphate, $Ca_3(PO_4)_2$, and sand, SiO_2, in an electric furnace?

10. What weight of phosphoric acid, 20% H_3PO_4, could be formed by dissolving in water the fumes from the combustion of 3 grams of phosphorus?

11. What volume of carbon monoxide, CO, would be formed during the manufacture of 50 kilograms of phosphorus from sand, SiO_2, carbon, calcium phosphate, $Ca_3(PO_4)_2$, and rock, in an electric furnace?

12. What weight of arsenic could be obtained from arsenic trioxide, As_2O_3, if an excess of the oxide is heated with 5 kilograms of carbon?

13. What per cent of antimony is present in an ore which contains 92% Sb_2S_3?

14. What weight of bismuth nitrate, $Bi(NO_3)_3$, would enter into hydrolysis in the formation of 30 grams of bismuth subnitrate, $BiONO_3$?

15. Theoretically, what volume of stibine, SbH_3, may be formed in the reaction between 0.3 gram of antimony chloride, $SbCl_3$, and nascent hydrogen?

16. What weight of arsenic may be recovered from 50 cc. of pure arsine, AsH_3?

17. The arsenic present in 0.4 gram of arsenic trioxide, As_2O_3, is converted to arsine, AsH_3, by nascent hydrogen. How many grams of zinc would be needed to produce sufficient hydrogen for the purpose?

18. What weight of phosphorus may be recovered from 130 cc. of phosphine, PH_3, which is decomposed by heating?

CHAPTER XVIII. SILICON AND BORON

The oxides of these elements are active acid anhydrides at high temperatures. In combination with the metals they form salts which generally are very stable.

1. What weight of pure silicon dioxide, SiO_2, would be required for the manufacture of 80.6 pounds of carborundum, CSi, by heating it with coke in an electric resistance furnace?

2. What amounts of silicon dioxide, SiO_2, and potassium carbonate, K_2CO_3, are required to produce 154.3 pounds of potassium silicate, K_2SiO_3, by fusion?

3. What is the per cent of boron, B, in crystallized borax, $Na_2B_4O_7 \cdot 10\,H_2O$?

4. How many pounds of carborundum, CSi, may be produced by the reaction between 100 pounds of silicon dioxide, SiO_2, and carbon, in the electric resistance furnace?

5. How many pounds of sodium silicate solution, 10% Na_2SiO_3, may be prepared from the product of the fusion of 10 pounds of anhydrous sodium carbonate, Na_2CO_3, with an excess of sand, SiO_2?

6. In the production of 2500 pounds of silicon, Si, by reduction of silicon dioxide, SiO_2, in an electric furnace, how many pounds of carbon would be needed, neglecting that used up in other reactions?

7. How many tons of carborundum, CSi, can be manufactured from 600 tons of silicon dioxide, SiO_2, in the electric furnace, if the process is 70% efficient?

8. 400 pounds of a mixture of silicon dioxide, 40%, and potassium hydroxide, 60%, is fused. What weight of water glass, K_2SiO_3, would be produced?

9. What weight of hydrogen fluoride, HF, would be used up in etching glass, while changing 40 grams of silicon dioxide to silicon tetrafluoride, SiF_4, and water?

10. What per cent of water of hydration is present in crystallized borax, $Na_2B_4O_7 \cdot 10\ H_2O$?

11. What weight of hydrochloric acid, 37% HCl, would be required to react with 20 pounds of crystallized borax, $Na_2B_4O_7 \cdot 10\ H_2O$, to form, as one of the products, boric acid, H_3BO_3?

12. What volume of hydrogen fluoride, HF, is needed to combine exactly with 5 grams of silicon dioxide, SiO_2?

13. What weight of hydrogen fluoride solution, 20% HF, is necessary to put into solution 12 grams of quartz, SiO_2?

CHAPTER XIX. THE METALS AND METALLURGY

The Metals

The metals occur in nature chiefly as oxides, carbonates, sulfides, and in the native condition. The general method for extracting metals from the oxides is by reduction, usually with carbon. Sulfides generally are roasted to form the oxides and are then reduced. Carbonates are generally decomposed by heat to form the oxides and are then reduced. In particular cases, reducing agents other than carbon may be required. Hydrogen or aluminum is commonly used in this way. In other cases, electrolysis may be needed to separate the metal from its compounds.

Displacement

If a piece of magnesium is placed in a solution of an acid, hydrogen is evolved, and the magnesium combines with the remainder of the acid to form magnesium chloride, $MgCl_2$. The reaction is:

$$Mg+2\,HCl \longrightarrow MgCl_2+H_2.$$

In the same way, when zinc is placed in a solution of copper sulfate, $CuSO_4$, copper is precipitated from the solution and a corresponding number of molecular weights of zinc enter into combination with the remaining radical of the copper sulfate:

$$Zn+CuSO_4 \longrightarrow ZnSO_4+Cu.$$

If a piece of copper were to be suspended in a solution of zinc sulfate, $ZnSO_4$ (the products of the above reaction), no change would result. A piece of copper, however, if suspended in a solution of silver nitrate, $AgNO_3$, would precipitate silver from the solution while a corresponding number of molecular weights of copper would combine with the remaining radical of the silver nitrate:

$$\mathrm{Cu} + 2\,\mathrm{AgNO_3} \longrightarrow \mathrm{Cu(NO_3)_2} + 2\,\mathrm{Ag}.$$

In these reactions the magnesium is said to displace the hydrogen from the acid, the zinc displaces the copper from the copper sulfate, and the copper displaces the silver from the silver nitrate. In displacement, one element gains a radical while another loses it; for example, the copper in the last reaction loses nothing, but gains two nitrate radicals, while the silver gains nothing, but loses these same two nitrate radicals.

Upon the results of a long series of experiments of this kind in which strips of metals are suspended in solutions of salts, one is enabled to build up readily the following series showing the relative displacing power of the various metals named, in which No. 1 has the greatest and No. 23 the least displacing power. This series is called a *Displacement Series.*

DISPLACEMENT SERIES

1. Potassium
2. Sodium
3. Magnesium
4. Aluminum
5. Manganese
6. Zinc
7. Chromium
8. Cadmium
9. Iron
10. Cobalt
11. Nickel
12. Tin
13. Lead
14. *Hydrogen*
15. Copper
16. Arsenic
17. Bismuth
18. Antimony
19. Mercury
20. Silver
21. Palladium
22. Platinum
23. Gold

The position of hydrogen in this series is especially important. One would assume from an inspection of the table that any metal preceding hydrogen in the table would release it from its salts (acids). This is true. One would also suspect that since both aluminum and iron release hydrogen from acids and since aluminum displaces iron from solutions of its salts, the aluminum should release hydrogen from acids more energetically than would iron. This is also true.

It will also be noticed from an inspection of this table, that none of the metals preceding hydrogen is found free in nature as an ordinary mineral, while those following hydrogen, though occurring in combination to some extent, may be found free.

The reason for this is that wherever the metals which precede hydrogen come in contact with water, which usually contains carbonic acid and other substances which contain hydrogen, they slowly displace the hydrogen and form compounds with the remaining elements. A table of chemical activity of the free metals is then seen to place them in the same order as the displacement series.

The earlier metals, *e.g.* potassium, sodium, iron, rust or tarnish much more readily in air than do the later ones. Those which follow copper, *e.g.* arsenic, mercury, silver, do not rust. On the other hand, we find that the oxides of the metals down to and including manganese, when heated in a stream of hydrogen, may give lower oxides, but are not completely reduced. The oxides of cadmium and succeeding metals are easily reduced. The oxides of the last five metals in this series, beginning with mercury, are decomposed by simply heating them. The activity of the metals toward other elements in addition to oxygen is similarly expressed by the arrangement of this table.

Ions and Displacement

From the above it can be seen that one kind of ion seems to enter more readily into combination than another kind of ion, and that in compounds one kind of ion seems to be more easily torn loose or displaced. This may be proved by a study of the electrolysis of their various salts. The electromotive force of the current required to release potassium from its compounds by electrolysis is greater than the electromotive force of the current required to release sodium.

This is true right down the list; the electromotive force of the current required to deposit any metal is less than that required for the preceding metal and more than that required for the one succeeding it. For this reason the same series which we have termed the *Displacement Series* may be termed the *Electromotive Series.*

1. What weight of silver, Ag, could be recovered from a solution containing 170 pounds of silver nitrate, $AgNO_3$, by displacement with zinc?

2. How much mercury, Hg, by weight could be obtained by roasting 1 kilogram of an ore containing 92.8% of cinnabar, HgS?

3. How many grams of copper, Cu, will be released from a solution of copper sulfate, $CuSO_4$, through displacement by the iron in a nail weighing 10.5 grams, if the nail is 98.8% Fe?

4. What volume of hydrogen would be used up in reducing to metallic copper, Cu, 50 grams of black copper oxide, CuO?

5. In reducing 5 pounds of ferric oxide, Fe_2O_3, by means of powdered aluminum, Al, what weight of aluminum would be used up?

6. What weight of zinc, Zn, would be needed to obtain by displacement 50 pounds of copper, Cu, from a solution of copper sulfate?

7. 70 pounds of silver, Ag, is recovered from a solution of silver nitrate, $AgNO_3$, by displacing the silver with zinc, Zn. How much zinc nitrate, $Zn(NO_3)_2$, was present in the solution at the close of the operation?

8. What weight of iron, Fe, could be released from ferric oxide, Fe_2O_3, through displacement by 20 pounds of powdered aluminum, Al?

9. How many tons of coke, 95% C., would be needed to reduce 100 tons of hematite, 98% Fe_2O_3, forming two parts of carbon monoxide, CO, to one part of carbon dioxide, CO_2?

10. What volume of hydrogen would be necessary to completely reduce one kilogram of strongly heated copper oxide, CuO?

11. What weight of zinc oxide, ZnO, would be left as the result of heating 5 kilograms of zinc sulfide, ZnS, in an unlimited supply of air?

12. How many tons of sulfuric acid, 70% H_2SO_4, could be manufactured from the sulfur in 80 tons of zinc blende ore, 95% ZnS?

13. What weight of powdered coal, 75% C, would be needed to reduce the oxide obtained by roasting 40 tons of zinc blende ore, 95% ZnS, in an unlimited air supply, if the carbon is all converted to carbon monoxide?

14. How many tons of carnallite, 95% $KCl \cdot MgCl_2 \cdot 6\,H_2O$, would be needed to furnish, by electrolysis, 1500 pounds of magnesium metal?

15. 60 tons of smithsonite, $ZnCO_3$, would lose how much in weight through roasting to the oxide, ZnO?

16. What per cent of copper, Cu, is present in chalcopyrite, 96% $CuFeS_2$?

17. What weight of silica, SiO_2, would be needed to produce, by reduction in an electric furnace, 50 kilograms of silicon, Si, if in the process 75% of the product is recovered?

18. 40 kilograms of copper oxide, CuO, is reduced by hydrogen to copper, Cu, and water. How much water would be needed to furnish by electrolysis the hydrogen required for this reduction? Assume that the process could be carried out with no waste of materials.

19. What weight of zinc blende, 30% ZnS, would be needed to produce 150 tons of sulfuric acid, 70% H_2SO_4?

20. What is the percentage of iron in chromite or chrome iron ore, $FeCr_2O_4$?

21. Calculate the per cent composition of iron alum,

$$NH_4Fe(SO_4)_2 \cdot 12\ H_2O.$$

CHAPTER XX. THE SODIUM FAMILY

Potassium, sodium, and lithium form hydroxides which are soluble in water. These metals invariably occupy the basic position in compounds. They occur always in combination, commonly as chlorides, and are usually extracted by electrolysis of the fused hydroxides. They cannot be exposed to air without change, as they react freely with the water and carbon dioxide.

1. What weight of sodium, Na, could be obtained by the electrolysis of 200 grams of fused sodium chloride, NaCl?

2. Assuming that there is no waste, how much common salt, NaCl, would be needed for the manufacture of 106 tons of sodium carbonate, Na_2CO_3?

3. How much pure sodium chloride, NaCl, would be used up in the production of 360 pounds of sodium hydroxide, NaOH, by electrolysis of sodium chloride solution?

4. How much sodium peroxide, Na_2O_2, would be needed to produce 110 liters of oxygen by its reaction with water?

5. What weight of potassium nitrate, KNO_3, could be formed in the reaction between potassium chloride, KCl, and 75 grams of sodium nitrate, $NaNO_3$?

6. What volume of hydrogen is evolved during the production of 1 kilogram of sodium, by electrolysis of fused sodium hydroxide, NaOH?

7. How many grams of water would be formed in the neutralization of 80 grams of sodium hydroxide, NaOH, by acetic acid, $HC_2H_3O_2$?

8. What volume of hydrogen chloride, HCl, will be formed by burning the hydrogen and the chlorine, evolved in the production of 1000 grams of sodium hydroxide, NaOH, by electrolysis of sodium chloride solution?

9. In the LeBlanc[1] process of making sodium carbonate, Na_2CO_3, how many pounds of hydrochloric acid, 40% HCl, could be made as a by-product during the manufacture of 1 ton of sodium carbonate?

10. How much sodium hydroxide, NaOH, would be formed as a by-product in the preparation of 50 liters of oxygen through the action of sodium peroxide, Na_2O_2, upon water?

11. It is desired to produce 150 grams of potassium nitrate, KNO_3, by the reaction between potassium chloride, KCl, and sodium nitrate, $NaNO_3$, in solution. What weights of potassium chloride and sodium nitrate should be used?

12. In the production of 60 kilograms of pure sodium through the electrolysis of melted sodium hydroxide, what weight of the hydroxide would be needed?

13. What weight of sodium could be obtained from 500 pounds of anhydrous sodium carbonate, Na_2CO_3, by reduction with carbon?

14. In the production of 40 liters of hydrogen, by the action of potassium on water, what weight of the metal will be used?

15. What weight of sodium chloride would be needed in the production of 75 pounds of sodium hydroxide, by electrolysis of its solution, if 85% of the hydroxide is recovered in the process?

16. In the lye process of obtaining sodium hydroxide, a solution of calcium hydroxide, $Ca(OH)_2$, is boiled with a solution of sodium carbonate, Na_2CO_3. What weight of sodium

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[1] A detailed description of this process will be found on page 133.

hydroxide could be obtained in this way from 300 pounds of sodium carbonate if the process is 90% efficient?

17. In the production of 5 liters of oxygen gas through the action of sodium peroxide, Na_2O_2, upon water, how much sodium hydroxide by weight would be formed?

18. What weight of sodium carbonate, Na_2CO_3, could be manufactured from 400 tons of sodium chloride, NaCl?

19. In the manufacture of sodium carbonate, Na_2CO_3, by the LeBlanc[1] process, what weight of calcium carbonate would be needed to convert the sodium sulfide, Na_2S, obtained from 90 tons of sodium chloride to sodium carbonate?

20. What weight of hydrochloric acid, 35% HCl, could be obtained as a by-product in the manufacture of 400 tons of sodium carbonate, Na_2CO_3, by the LeBlanc[1] process?

21. During the manufacture of 250 tons of sodium carbonate, by the Solvay[2] process, what weight of sodium bicarbonate, $NaHCO_3$, would be produced as a step in the process?

22. What weight of potassium nitrate, KNO_3, can be obtained by heating a solution containing 1 kilogram of sodium nitrate, 94% $NaNO_3$, with a solution of potassium chloride, KCl, if 93% of the potassium nitrate is recovered in the process?

23. Hypo, or sodium thiosulfate, $Na_2S_2O_3 \cdot 5\ H_2O$, can be prepared by heating a solution of sodium sulfite, Na_2SO_3, with sulfur. What weight of hypo can be made in this way from 360 pounds of sodium sulfite?

24. What is the per cent composition of hypo, or sodium thiosulfate, $Na_2S_2O_3 \cdot 5\ H_2O$?

25. In order to test the purity of a sample of anhydrous borax, $Na_2B_4O_7$, 1 gram of the sample was dissolved in 100 cc.

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[1] A detailed description of this process will be found on page 133.

[2] For a detailed description of this process, see page 134.

of water and the solution was titrated with normal hydrochloric acid solution, using methyl orange as an end-point indicator. 9 cc. of the normal acid solution were required. What was the per cent purity of the borax?

$$Na_2B_4O_7 + 2\ HCl + 5\ H_2O \longrightarrow 2\ NaCl + 4\ H_3BO_3.$$

26. A student is given 400 g. of washing soda crystals, $Na_2B_4O_7 \cdot 10\ H_2O$, with the object of making caustic soda, NaOH, from it. What weight of product would he get, assuming that there is no waste in the process?

27. A dealer contracts to sell 200 lbs. of crystallized washing soda, $Na_2B_4O_7 \cdot 10\ H_2O$, at 8 cents per lb. What would be his loss if the soda which he sells has lost one fourth of its water of hydration?

CHAPTER XXI. GLYCERIN AND EXPLOSIVES

A substance, in order to make a satisfactory explosive, must produce a very large volume of gas at a high temperature and within a very short space of time. Explosives may be of the type which decomposes under shock, like nitroglycerin, the explosion of which is instantaneous and develops full pressure at once, or they may be of the type, such as smokeless gunpowder, the burning of which may be somewhat controlled. In the latter case the pressure develops as the combustion proceeds. Many of the high explosives are not sensitive to ordinary shock but require a detonator to set them off. T. N. T., tri-nitro-toluol, is a good example of the type of explosive requiring a detonator.

1. Calculate the percentage proportions in which the constituents of black gunpowder should be mixed in its preparation. The equation representing the principal reaction in its explosion is:

$$2\ KNO_3 + 3\ C + S \longrightarrow K_2S + 3\ CO_2 + N_2.$$

2. 100 pounds of glycerin, $C_3H_5(OH)_3$, is nitrated to produce nitroglycerin, $C_3H_5(NO_3)_3$. What weight of nitroglycerin is formed?

3. What weight of water would be formed in making 800 pounds of nitroglycerin, $C_3H_5(NO_3)_3$, by nitrating glycerin, $C_3H_5(OH)_3$, in a mixture of concentrated sulfuric and nitric acids?

4. The principal change upon the ignition of ordinary black gunpowder is indicated in the equation:

$$2\,KNO_3 + 3\,C + S \longrightarrow K_2S + 3\,CO_2 + N_2.$$

What theoretical weights of potassium nitrate, charcoal, and sulfur would be needed to make up 150 pounds of gunpowder?

5. What volume of gas, carbon dioxide and nitrogen, would be released by the explosion of 1 kilogram of black gunpowder, gas volume measured at standard temperature and pressure?

6. The changes taking place in the decomposition or explosion of nitroglycerin, $C_3H_5(NO_3)_3$, are represented, in a general way, by the equation:

$$4\,C_3H_5(NO_3)_3 \longrightarrow 12\,CO_2 + 6\,N_2 + 10\,H_2O + O_2.$$

What total volume of gases would be released by the explosion of 50 grams of nitroglycerin, gases measured at standard temperature and pressure?

7. (*a*) The exploding temperature of nitroglycerin is approximately 1850° C. What volume would the gases, given off in the explosion of 1 kilogram of nitroglycerin, occupy if they were at the exploding temperature but at standard air pressure? The explosion of nitroglycerin is represented, in a general way, by the equation:

$$4\,C_3H_5(NO_3)_3 \longrightarrow 12\,CO_2 + 6\,N_2 + 10\,H_2O + O_2.$$

(*b*) What pressure would be exerted by the exploding gases if they were confined within a space of 2 liters at the exploding temperature?

8. What weight of glycerin would be needed in the preparation of 5 tons of nitroglycerin, by nitration?

CHAPTER XXII. THE CALCIUM FAMILY

The metals of the calcium group form hydroxides which are slightly soluble in water. They occur commonly as carbonates which may be decomposed to oxides by heating. Their oxides are practically infusible and are stable at high temperatures. Nearly all hardness in water is due to the presence of soluble compounds of these metals.

1. How many liters of carbon dioxide would be produced by the reaction between 200 grams of calcium carbonate, $CaCO_3$, and hydrochloric acid?

2. How many tons of limestone or calcium carbonate, $CaCO_3$, would be needed in order to produce, through strong heating, 560 tons of quicklime or calcium oxide, CaO?

3. What quantity of plaster of Paris, $(CaSO_4)_2 \cdot H_2O$, can be obtained from 1032 pounds of gypsum, $CaSO_4 \cdot 2\,H_2O$, by heating the gypsum at a uniform temperature of about 125° C. until the change is complete?

4. How many pounds of sulfuric acid, 98% H_2SO_4, would be needed to convert 930 pounds of normal calcium phosphate, $Ca_3(PO_4)_2$, to the monocalcium acid phosphate, $CaH_4(PO_4)_2$, used for fertilizer?

5. How many grams of calcium carbonate, $CaCO_3$, are required for the production of 100 liters of carbon dioxide by its reaction with dilute nitric acid?

6. 100 tons of limestone, 98% $CaCO_3$, is burned to quicklime, CaO, by heating to a red heat, 900° C., until the carbonate is completely decomposed. How many cubic feet of carbon dioxide are produced? 1 cubic foot of carbon dioxide weighs 0.0125 pound.

7. How much gypsum, $CaSO_4 \cdot 2\,H_2O$, is required to yield 20 tons of plaster of Paris, $(CaSO_4)_2 \cdot H_2O$, by uniformly heating the gypsum at a temperature of about 125° C. until the change is complete?

8. How many grams of soap, having the formula $NaC_{18}H_{35}O_2$, are required to soften 1 kilogram of water which contains 15 parts by weight of calcium hydrogen carbonate, $CaH_2(CO_3)_2$, per million?

9. Plaster of Paris has the formula, $(CaSO_4)_2 \cdot H_2O$. What is the percentage of water present in the molecule?

10. A mass of crystallized barium chloride which had a weight of 50 grams was heated until no further change in weight took place, when it was found to have lost 7.375 grams of water of hydration. From these data calculate the percentage of water of hydration in the crystal and also the complete formula.

11. What weight of limestone, $CaCO_3$, would be required to yield 850 tons of quicklime, CaO, when acted upon by a red heat, 900° C.?

12. A quantity of water containing 0.4 gram of calcium bicarbonate, $CaH_2(CO_3)_2$, hardness to 1 liter is softened by the addition of 60 grams of slaked lime, $Ca(OH)_2$. How many liters of water were there?

13. A mass of mortar containing 75 grams of slaked lime or calcium hydroxide, $Ca(OH)_2$, slowly hardens by having the calcium hydroxide converted to calcium carbonate, $CaCO_3$, through the absorption of carbon dioxide from the air. What volume of carbon dioxide is absorbed in the process?

14. The French or ordinary laboratory method of estimating 1 degree of calcium carbonate hardness, or its equivalent of

other salt, means 0.01 gram per liter. What weight of a soap of the formula, $NaC_{16}H_{31}O_2$, sodium palmitate, would be needed to soften 50 liters of a water possessing a calcium carbonate hardness of 28 degrees? 1 degree of hardness means a precipitation of the quantity of material stated.

15. What weight of calcined or dehydrated washing soda, Na_2CO_3, is used up in softening 80 gallons of well water possessing a calcium sulfate, $CaSO_4$, hardness of 35 degrees, French?

16. What weight of bleaching powder, $CaOCl_2$, would be formed by the reaction between 450 liters of chlorine and freshly slaked lime, $Ca(OH)_2$?

CHAPTER XXIII. THE MAGNESIUM FAMILY

The calcium and magnesium families constitute one group and are, in general, similar in properties.

1. How many grams of magnesia, MgO, could be prepared by heating 843 grams of magnesium carbonate, $MgCO_3$, until the reaction is complete?

2. What is the per cent of magnesium, Mg, in Epsom salt, $MgSO_4 \cdot 7\,H_2O$?

3. How many tons of zinc oxide, ZnO, could be produced by the roasting in air of 974 tons of zinc blende ore, ZnS?

4. What weight of zinc sulfide, ZnS, could be formed by passing 10 liters of hydrogen sulfide, H_2S, over an excess of heated zinc oxide, ZnO?

5. How many pounds of zinc metal, Zn, may be produced by roasting in air 100 pounds of zinc blende, ZnS, and then reducing the oxide formed by means of carbon, allowing for a 3% waste?

6. How many pounds of magnesium carbonate, $MgCO_3$, are required to produce 20 pounds of magnesia, MgO, by heating until the carbonate is completely decomposed?

7. What weight of soap with the formula, $KC_{18}H_{35}O_2$, will be used up in softening 1 kilogram of water containing 0.05% of magnesium chloride, $MgCl_2$, in solution?

8. How much zinc might be produced from 5 tons of zinc blende, ZnS, allowing a waste of 4%?

9. In analyzing a certain paint, it was found to contain 30% zinc metal by weight. What weight of zinc oxide, ZnO, would be required to make 100 pounds of this paint?

10. Crystallized magnesium sulfate has the formula, $MgSO_4 \cdot 7\ H_2O$. How much would 100 pounds of the crystals lose in weight upon conversion to the anhydrous magnesium sulfate?

11. What is the per cent composition of zinc sulfate, $ZnSO_4 \cdot 7\ H_2O$?

12. What weight of sulfuric acid, 70% H_2SO_4, may be manufactured from the sulfur dioxide which is given off during the roasting of 25 tons of zinc blende ore, 25% ZnS?

CHAPTER XXIV. ALUMINUM

Aluminum is the lightest metal in common use. Pure aluminum is practically unaltered in air. When the foil or powder is strongly heated it burns with a very brilliant light which resembles that of magnesium. The metal easily releases hydrogen from hydrochloric acid and forms aluminum chloride. It also releases hydrogen from sodium and potassium hydroxides, forming aluminates of the respective metals.

On account of its great affinity for oxygen it is often used to extract metals from their oxides. Many metals like chromium, uranium, and manganese are now reduced in this manner. The process which was named *aluminothermy* by its inventor, Goldschmidt, is now called the Thermit process.

1. In the reaction between sodium hydroxide and 54 grams of aluminum, Al, to form sodium aluminate, Na_3AlO_3, how many liters of hydrogen would be released?

2. How many pounds of aluminum, Al, would be needed to obtain 168 pounds of pure iron, Fe, by the thermit process?

3. How many pounds of aluminum sulfate, $Al_2(SO_4)_3$, would be required to remove 486 pounds of calcium bicarbonate, $CaH_2(CO_3)_2$, from water, by precipitation? The probable reactions are as follows:

$$Al_2(SO_4)_3 + 3\ CaH_2(CO_3)_2 + 6\ H_2O \longrightarrow 6\ H_2CO_3 + 3\ CaSO_4 + 2\ Al(OH)_3.$$

4. What per cent of aluminum, Al, is present in potassium alum, $AlK(SO_4)_2 \cdot 12\ H_2O$?

5. How much hydrochloric acid, 40% HCl, is required to put 80 grams of aluminum into solution as aluminum chloride, $AlCl_3$?

6. It is desired to make 1 ounce of flashlight powder by mixing aluminum powder, Al, and potassium chlorate, $KClO_3$. What quantity of each should be used?

7. In the manufacture of 1000 pounds of alum baking powder, what weights of anhydrous potassium alum, $AlK(SO_4)_2$, and of baking soda, $NaHCO_3$, should be used, if the baking powder formula includes 10% corn starch to prevent the caking of the powder?

8. What volume of carbon dioxide, measured at a pressure of 760 mm. of mercury and a baking temperature of 120° C., would be produced by the interaction of aluminum potassium sulfate, $AlK(SO_4)_2$, and baking soda, $NaHCO_3$, in 10 grams of a baking powder made up of these two substances in the proper proportions to combine completely with each other? 10% of the baking powder is composed of corn starch to prevent its caking or reacting from the collection of moisture.

9. The ruby is crystallized aluminum oxide, Al_2O_3, with a trace of coloring material. What weight of aluminum, Al, is present in a ruby weighing 0.15 gram?

10. 3 grams of sodium hydroxide, NaOH, in solution is boiled in an aluminum saucepan. Provided that all of the sodium hydroxide reacts with the aluminum, what is the loss in weight of the saucepan?

11. 2.2 grams of potassium hydroxide, KOH, in solution is added to a precipitate of aluminum hydroxide, $Al(OH)_3$. How much aluminum chloride, $AlCl_3$, may be prepared from the resulting solution if an excess of hydrochloric acid is added to it?

12. What weight of baking soda, $NaHCO_3$, should be added to 30 grams of anhydrous potassium alum, $AlK(SO_4)_2$, to react completely with it in solution?

13. Kaolin has the formula, $H_4Al_2Si_2O_9$. What is its per cent composition?

14. Feldspar has the formula, $KAlSi_3O_8$. What is its per cent composition?

15. How many pounds of aluminum, Al, are required to produce 12 pounds of chromium, Cr, from chromic oxide, Cr_2O_3?

16. How much aluminum oxide, Al_2O_3, is produced as a slag in obtaining 16 grams of manganese from manganese dioxide, MnO_2, by the thermit process?

17. 0.2 gram of magnesium, Mg, is suspended in a solution of aluminum chloride, $AlCl_3$. What weight of aluminum will be displaced by the magnesium?

18. 3.0 grams of aluminum, Al, are placed in a solution of zinc nitrate, $Zn(NO_3)_2$. What weight of zinc, Zn, will be displaced by the aluminum?

CHAPTER XXV. THE IRON FAMILY

Iron occurs chiefly as ferric oxide in which the iron has a valence of three. Carbonates and sulfides also occur in abundance. Iron is generally obtained through the reduction of its oxides by carbon. Iron, with its main product, steel, is the most important of all metals economically. Nickel occurs principally as complex sulfides and silicates, from which it is extracted by special processes. Cobalt, as an element, has no well-developed use. All three of these metals possess magnetic properties.

1. A mass of iron sulfide, FeS, contains 20 pounds of pure iron, Fe, in combination. What is the weight of the iron sulfide?

2. How much iron, 94% Fe, could be obtained from 1 ton of hematite, 98% Fe_2O_3?

3. How much coke, 95% C, is needed for the actual reduction of 100 tons of hematite, Fe_2O_3, in the blast furnace, to metallic iron? Assume that the gases formed are: carbon monoxide, 2 parts, and carbon dioxide, 1 part by volume.

4. What weight of coke, 96% C, would be used up in the reduction of 300 tons of hematite ore, Fe_2O_3, assuming that the products of the reduction are iron and 2 parts of carbon monoxide, CO, to 1 part of carbon dioxide, CO_2, by volume?

5. What weight of ferric hydroxide, $Fe(OH)_3$, could be formed by adding to a solution of ferric chloride, $FeCl_3$, a solution containing 50 grams of sodium hydroxide?

6. What volume of sulfur dioxide, SO_2, would result from the complete oxidation of 1 ton of pure iron pyrites, FeS_2? 1 pound is equal to 453.6 grams.

7. Calculate the per cent of water of hydration in crystallized green vitriol or ferrous sulfate, $FeSO_4 \cdot 7\,H_2O$.

8. What weight of ferric hydroxide, $Fe(OH)_3$, will be precipitated from 50 cc. of a normal solution of ferric chloride, $FeCl_3$, by adding an excess of ammonium hydroxide, NH_4OH?

9. How many pounds of ferric oxide, Fe_2O_3, are required to produce 20 pounds of crystallized ferric nitrate, $Fe(NO_3)_3 \cdot 6\,H_2O$, by the action of dilute nitric acid?

10. What weight of sulfuric acid is required to convert 20 grams of ferrous sulfate, $FeSO_4$, to ferric sulfate, $Fe_2(SO_4)_3$, in the presence of an oxidizing agent?

11. What weight of zinc is required to produce enough hydrogen, in the nascent condition, to reduce 8.6 grams of ferric chloride, $FeCl_3$, to ferrous chloride, $FeCl_2$?

12. A 1-gram sample of steel is put into solution with an acid, heated to dryness, redissolved in an acid, and then filtered and washed. The residue consists of carbon and silica, SiO_2. This is heated until all of the carbon is burned out as carbon dioxide, and the residue of silica, SiO_2, found to have a weight of 0.0032 gram. What is the per cent of silicon in the steel?

13. 5 tons of molten steel in a ladle has a carbon content of 0.38%. How much pig iron with a carbon content of 4.0% must be added to raise the carbon in the steel to 0.50%?

14. What is the per cent of nickel in nickel carbonyl, $Ni(CO)_4$?

15. 2.5 grams of zinc, Zn, is placed in a solution of ferrous sulfate, $FeSO_4$. What weight of iron will be displaced by the zinc?

16. 2.9 grams of iron, Fe, is placed in a solution of copper sulfate, $CuSO_4$. What weight of ferrous sulfate, $FeSO_4$, will be produced?

17. What weight of iron, Fe, is required to recover the silver, Ag, in the form of metal, from 20 grams of a 10% solution of silver nitrate, $AgNO_3$?

18. 10 grams of oxide of iron is heated in hydrogen and the resulting iron weighs 7 grams. What is the formula of the oxide?

19. 4 tons of hematite (considered as pure Fe_2O_3) is smelted in a blast furnace with coke for the production of iron. If the coke contains 10% of ash, calculate the quantity of coke needed for the reduction. Assume that the gaseous products will contain 2 parts by volume of carbon monoxide to 1 part of carbon dioxide.

CHAPTER XXVI. COPPER, MERCURY, SILVER

These metals occur native and also as sulfides. The extraction from the sulfide is essentially a process of roasting and reduction, although the varying conditions and mixtures in the ores have made necessary many special modes of treatment.

Since these metals occur below hydrogen in the Displacement Series of metals, they are not much affected by acids in dilute condition, but react readily under oxidizing conditions. For example, such oxidizing agents as concentrated nitric acid and hot concentrated sulfuric acid readily oxidize the metals and then convert them to salts of the acids while the acids themselves are reduced and decomposed.

The metals also combine directly with some elements. The combination of copper and chlorine is a good illustration of this.

1. Chalcopyrite, a common ore of copper, has the formula, $Cu_2S \cdot Fe_2S_3$. How many tons of copper, Cu, could be obtained from 800 tons of this ore if 85% of the ore corresponds to this formula?

2. What weight of blue vitriol, $CuSO_4 \cdot 5\,H_2O$, could be manufactured from 400 pounds of scrap copper?

3. In one process of refining, silver is recovered from a solution of silver sulfate, Ag_2SO_4, by displacing the silver with copper. What weight of blue vitriol or copper sulfate will be formed as a by-product while obtaining 60 pounds of silver in this way?

4. How many grams of silver nitrate, $AgNO_3$, could be prepared by the action of 150 pounds of nitric acid, 60% HNO_3, upon silver, Ag?

5. What volume of sulfur dioxide would be formed as a by-product while obtaining 400 pounds of mercury, Hg, through the roasting of cinnabar, HgS, in contact with air?

6. How much silver nitrate, $AgNO_3$, by weight, can be made from a dime weighing 2.45 grams and containing 10% of alloy?

7. What weight of copper, Cu, may be obtained through the reduction of 30.2 grams of copper oxide, CuO, by heating it with charcoal?

8. How many grams of cupric nitrate, $Cu(NO_3)_2$, may be produced by the reaction between nitric acid and 16 grams of copper, Cu?

9. What is the per cent of water of hydration in crystallized copper sulfate, $CuSO_4 \cdot 5\ H_2O$?

10. What is the per cent of copper, Cu, in crystallized cupric chloride, $CuCl_2 \cdot 2\ H_2O$?

11. How many pounds of mercury, Hg, might be obtained from 100 pounds of crude cinnabar, which is 95.5% HgS, by roasting the ore?

12. How many grams of mercuric chloride, $HgCl_2$, may be produced by the action of aqua regia upon 3.5 grams of mercury, Hg?

13. How many grams of mercurous chloride, HgCl, may be precipitated from a solution of mercurous nitrate, $HgNO_3$, by means of a solution containing 10 grams of sodium chloride, NaCl?

14. How many pounds of silver, Ag, may be recovered from 1 ton of silver ore, 30% silver sulfide, or Ag_2S?

15. How many pounds of silver, Ag, must be dissolved in nitric acid to produce 35 pounds of silver nitrate, $AgNO_3$?

16. What weight of salt, NaCl, must be put into solution to precipitate 0.362 gram of silver chloride, AgCl, from a solution of silver nitrate, $AgNO_3$?

17. 20 grams of cupric nitrate, $Cu(NO_3)_2 \cdot 6\ H_2O$, is heated until completely decomposed. The copper is recovered as the oxide, CuO. What weight of copper oxide is obtained?

18. When a mixture of mercuric sulfate, $HgSO_4$, and sodium chloride, NaCl, is sublimed, mercuric chloride, $HgCl_2$, is produced. How many grams of mercuric chloride may be obtained by subliming 15.2 grams of sodium chloride with mercuric sulfate?

19. How many tons of anhydrous copper sulfate, $CuSO_4$, are required to produce 40 tons of copper, Cu, by electrolysis of its solution?

20. A certain current of electricity, flowing through a solution of silver nitrate, $AgNO_3$, deposits 0.16 gram of silver, Ag, on the cathode per hour. How many grams of hydrogen could be liberated from an acid in the same time by an equal flow of current?

21. A silver coin contains alloy, 10% by weight. How many grams of silver chloride, AgCl, could be made from a dime weighing 2.45 grams?

22. 3 grams of silver, Ag, is heated with sulfur to produce silver sulfide, Ag_2S. How many grams of sulfur enter into the reaction?

23. How many grams of copper would be deposited in one hour from a solution of cupric sulfate, $CuSO_4$, by such a current of electricity as would in the same time release 16 grams of oxygen from water?

CHAPTER XXVII. TIN AND LEAD

Tin and lead are usually found in combination with oxygen or with sulfur. Their metallurgy is essentially a matter of reduction of the oxide.

1. What weight of tin, Sn, could be obtained from 700 kilos of stannic oxide, SnO_2, through reduction with coal in a reverberatory furnace?

2. The Dutch process of making white lead may be represented by the equation:

$$6\,Pb + 4\,CO_2 + 2\,H_2O + 3\,O_2 \longrightarrow 2(2\,PbCO_3 \cdot Pb(OH)_2).$$

What weight of lead would be used up in making 100 pounds of white lead?

3. The purity of white lead is usually determined by observing the weight of carbon dioxide given off when it is treated with an acid. On the supposition that the formula, $(PbCO_3)_2 \cdot Pb(OH)_2$, is the correct formula, how nearly pure is a sample if 1 gram yields 102.1 milligrams of carbon dioxide when it is treated with an acid?

4. What weight of concentrated sulfuric acid, 98% H_2SO_4, can be manufactured from the sulfur in one ton of galena, PbS?

5. Red lead or American vermilion, Pb_3O_4, can be made by heating lead or lead oxide, PbO, in the air, oxygen being absorbed in the operation. This is used as a drier in paints on account of its oxidizing character. What weight of oxygen is absorbed in converting 1 ton of lead oxide, PbO, to red lead, Pb_3O_4?

6. Metallic lead has a specific gravity of 11.35. How large a block of lead might be obtained by the reduction of 80 grams of litharge, PbO?

7. The amount of lead present in a substance is sometimes determined by putting the substance into solution with nitric acid and then precipitating the lead in the form of lead sulfate. In one such case a 3-gram sample yielded 2.5 grams of lead sulfate, $PbSO_4$. What was the per cent of lead in the sample?

8. A shipment of 10 tons of cassiterite or tinstone is found to analyze 80% SnO_2. Provided that 97% of the tin is recovered as metal, how many tons of tin would this shipment yield?

9. What weight of stannic chloride, $SnCl_4$, may be produced by oxidizing a solution of 20 grams of stannous chloride, $SnCl_2$, in hydrochloric acid, by means of nitric acid?

10. What weight of lead sulfide, PbS, will be formed by the action of hydrogen sulfide, H_2S, upon 2.5 grams of lead nitrate, $Pb(NO_3)_2$?

11. What is the per cent composition of lead sulfate, $PbSO_4$?

CHAPTER XXVIII. RADIUM AND THORIUM

Study of the behavior of these elements since the original investigations by Becquerel and others has opened a tremendous field of investigation into the structure of the atom. It is safe to say that our conception of the atom as the unit of chemical change has not been changed by any data on atomic structure. We have, however, gained new and more complete appreciation of the properties of the elements as related to their atomic weights and structure.

1. Radium chloride, $RaCl_2$, is worth $100,000 per gram, the value being estimated upon the amount of radium present. What would be the value of 10 milligrams of radium sulfate, $RaSO_4$?

2. A piece of cotton tubing is saturated with 3.0 grams of a 10% solution of thorium nitrate, $Th(NO_3)_4$. When the tubing — Welsbach gas mantle — is burned, what weight of thorium oxide, ThO_2, will remain in the ash of the mantle?

CHAPTER XXIX. MANGANESE AND CHROMIUM

These metals are most useful in the form of alloys, particularly with steel, to which they impart certain very valuable properties. They are usually found in combination with oxygen, from which they may be separated by means of reduction with aluminum (Thermit process).

1. Manganese, Mn, is obtained by igniting a mixture of pyrolusite, MnO_2, and aluminum powder, Al. What weight of aluminum would be required to release the manganese from 1 ton of this ore, 60% MnO_2?

2. Calculate the per cent of chlorine present in crystallized manganous chloride, $MnCl_2 \cdot 4\,H_2O$.

3. Potassium chromate, K_2CrO_4, is prepared by roasting chromite, $FeO \cdot Cr_2O_3$, with potash, K_2CO_3, and lime, CaO, according to the equation:

$$4\,FeO \cdot Cr_2O_3 + 8\,K_2CO_3 + 7\,O_2 \longrightarrow 2\,Fe_2O_3 + 8\,CO_2 + 8\,K_2CrO_4$$

The potassium chromate is dissolved out of the residue. What weight of potassium chromate can be obtained in this way from 1 ton of the chromite, if the process is 75% efficient?

4. What weight of manganese, Mn, could be produced by the reduction of 10 grams of manganese dioxide, MnO_2, in the thermit process, by means of aluminum, Al?

5. 30 grams of manganese dioxide, MnO_2, is converted into potassium manganate, K_2MnO_4, by fusion with potassium hydroxide and an oxidizing agent as potassium nitrate. How much potassium manganate may thus be prepared?

6. A concentrated solution of potassium manganate, K_2MnO_4, is diluted with a considerable amount of water, when the following reaction takes place:

$$3\ K_2MnO_4 + 2\ H_2O \longrightarrow 4\ KOH + MnO_2 + 2\ KMnO_4.$$

How much potassium permanganate, $KMnO_4$, could be obtained in this way from 12 grams of potassium manganate?

7. 26 grams of chromic oxide, Cr_2O_3, is mixed with sufficient aluminum and the mixture is ignited by means of magnesium (thermit reaction). How many grams of aluminum are required to complete the reduction of the chromic oxide to chromium?

8. 130 pounds of chromite, $FeO \cdot Cr_2O_3$, is roasted with potash, K_2CO_3, and lime, CaO, when the following reaction takes place:

$$4\ FeO \cdot Cr_2O_3 + 8\ K_2CO_3 + 7\ O_2 \longrightarrow 2\ Fe_2O_3 + 8\ CO_2 + 8\ K_2CrO_4.$$

What weight of potassium chromate, K_2CrO_4, would this yield?

9. 56 grams of potassium chromate, K_2CrO_4, in solution, is treated with sulfuric acid, and potassium dichromate, $K_2Cr_2O_7$, is formed as follows:

$$2\ K_2CrO_4 + H_2SO_4 \longrightarrow H_2O + K_2SO_4 + K_2Cr_2O_7.$$

What weight of potassium dichromate will be formed?

10. When potassium hydroxide, KOH, is added to a solution of potassium dichromate, $K_2Cr_2O_7$, the following reaction takes place:

$$K_2Cr_2O_7 + 2\ KOH \longrightarrow H_2O + 2\ K_2CrO_4.$$

If it is desired to produce 16 grams of potassium chromate, K_2CrO_4, in this way, what weight of potassium dichromate must be used?

11. What per cent of chromium, Cr, is present in chrome alum, $KCr(SO_4)_2 \cdot 12\ H_2O$?

12. 20 cc. of a normal solution of chromium sulfate, $Cr_2(SO_4)_3$, is treated with 20 cc. of a normal solution of sodium hydroxide, NaOH. What weight of chromic hydroxide, $Cr(OH)_3$, is precipitated?

13. What is the per cent composition of potassium permanganate, $KMnO_4$?

14. What weight of potassium chromate, K_2CrO_4, can be prepared from potassium hydroxide and 300 grams of potassium dichromate, $K_2Cr_2O_7$?

CHAPTER XXX. PLATINUM AND GOLD

The usefulness of these metals depends primarily upon their chemical inactivity. They resist the action of all ordinary acids, but a mixture of concentrated hydrochloric and nitric acids, known as Aqua Regia, will convert them into soluble chlorine compounds by the action of nascent chlorine upon them.

1. What weight of platinum, Pt, will be deposited in a mass of asbestos fiber which has been soaked with 3 cc. of a 5% solution of chloroplatinic acid, H_2PtCl_6, and ignited over a bunsen flame?

2. At $4.50 per gram, what is the value of the platinum, Pt, present in 5 grams of a 10% solution of chloroplatinic acid, H_2PtCl_6?

3. Gold plating may be accomplished by the electrolysis of a solution of sodium aurocyanide, $NaAu(CN)_2$. To deposit a plating of gold, Au, weighing 0.2 gram, what weight of sodium aurocyanide will be decomposed?

4. 14-karat gold contains 14 parts of gold in 24 parts of material, the other 10 parts being alloy. What weight of auric chloride, $AuCl_3$, could be produced by treating a 14-karat gold ring, weighing 20 grams, with aqua regia, until the gold is entirely acted upon by the nascent chlorine?

CHAPTER XXXI. ORGANIC COMPOUNDS OF CARBON

Carbon is known to form over 200,000 definite compounds. Most of these are complex in nature, a great series very often being represented by the same empirical formula, so that a formula, in order to be of value to the chemist, must often be graphical or structural.

In other words, the formula must represent the elements present in the compound and also the manner in which the atoms are linked together to form the molecule. In all such compounds the carbon is *quadrivalent,* which means that its valence is four. This fact, also, can often be recognized only through a graphic representation of the molecular structure.

1. What is the per cent composition of cane sugar, $C_{12}H_{22}O_{11}$?

2. Cane sugar has the formula, $C_{12}H_{22}O_{11}$. What weight of pure carbon, C, may be produced by removing the hydrogen and oxygen from 30 grams of cane sugar by means of concentrated sulfuric acid?

3. What weight of glycerin, $C_3H_5(OH)_3$, would be formed as a by-product in the manufacture of 1 ton of a soap of the formula, $NaC_{18}H_{35}O_2$, sodium stearate?

4. What weight of a soft soap, composed of potassium palmitate, $KC_{16}H_{31}O_2$, can be made from a fat consisting of 300 pounds of palmitin, $C_3H_5(C_{16}H_{31}O_2)_3$?

5. 100 pounds of olein, $C_3H_5(C_{18}H_{33}O_2)_3$, is converted to stearin, $C_3H_5(C_{18}H_{35}O_2)_3$, by causing it to combine with hydrogen through the aid of powdered nickel as a catalyzer. How many pounds of stearin would be formed?

6. What volume of ether of a specific gravity of 0.7183 could be obtained by distilling at 35° C. 1 liter of ethyl alcohol having a specific gravity of 0.8 with concentrated sulfuric acid, neglecting the water present in the alcohol and in the product?

7. Ether, $(C_2H_5)_2O$, is prepared by distilling at 35° C. a mixture of ethyl alcohol, C_2H_5OH, and concentrated sulfuric acid which serves as a dehydrating agent. What weight of ether could be made in this way from one kilogram of alcohol, 95% C_2H_5OH?

CHAPTER XXXII. SPECIFIC HEAT

Substances other than water require different characteristic amounts of heat to change their temperatures. The ratio of heat required to raise the temperature of one gram of any substance through one degree C. to the amount of heat required to raise the temperature of one gram of water through one degree C. is called the *specific heat* of the substance.

In other words, the amount of heat required to raise the temperature of one gram of any substance through one degree C. is called the specific heat of the substance. This is usually determined by measuring the amount of heat given out by a weighed amount of the substance in cooling through a measured number of degrees C.

Example: 100 grams of lead shot at 100° C. is mixed with 100 grams of water at 15° C. The final temperature of the mixture is 17.5956° C. What is the specific heat of the lead?

Solution: The lead has cooled through a temperature range of 100° minus 17.5956° C., which is equal to 82.4044° C. The rise in temperature of the water is 17.5956° C. minus 15° C., or 2.5956° C. 2.5956×100 (grams of water) gives us 259.56 calories absorbed by the water. It is evident that if no heat escapes from the system, the heat absorbed by the water must have come from the lead in cooling.

Now dividing 259.56 by 100 (the number of grams of lead), we get 2.5956 calories given out by each gram of lead in cooling through 82.4044° C. Dividing 2.5956 by 82.4044° C. we get *0.0315*, which is the number of calories of heat given out by one gram of lead in cooling through one degree C. This is the specific heat of lead.

Du Long and Petit's Law

It has been found that when the atomic weight of an element is multiplied by the specific heat of the element, the product has nearly a constant value, regardless of what element is chosen. This constant is usually taken as 6.25.

Accordingly, if the specific heat of an element is determined and the constant 6.25 divided by the specific heat, the resulting value will be close enough to the true atomic weight to serve as a check on other methods of determining the atomic weight.

If a serious blunder is made in determining the atomic weight of an element, this principle, which is called the law of Du Long and Petit, will indicate the blunder at once.

Example: The specific heat of silver is 0.0559. Dividing the constant 6.25 by this value, we get 111.8, which nearly approaches the true atomic weight for silver, 107.88. If by some other method the atomic weight for silver had been found to be 215, the error would be indicated at once.

1. The specific heat of platinum is 0.0323. Calculate from this its approximate atomic weight.

2. The specific heat of copper is 0.0923. Calculate from this its approximate atomic weight.

3. The atomic weight of aluminum is 27.1. Calculate from this its approximate specific heat.

4. The atomic weight of sulfur is 32.06. Calculate from this its approximate specific heat.

5. The specific heat of calcium is 0.170. Calculate from this its approximate atomic weight.

6. The specific heat of antimony is 0.052. Find its approximate atomic weight.

7. The atomic weight of lithium is 6.94. Calculate its approximate specific heat.

8. The atomic weight of potassium is 39.1. Calculate its approximate specific heat.

9. The specific heat of uranium is 0.0276. From this find its approximate atomic weight.

10. What is the approximate specific heat of lead if its atomic weight is 207.1?

11. The per cent of gold in a compound of gold and chlorine is 64.958%. The atomic weight of chlorine is 35.46 and the specific heat of gold is 0.032. What is the formula of the chloride?

APPENDIX

The Haber Process

This is a process for making ammonia by direct combination of nitrogen and hydrogen. High pressure (185 to 200 atmospheres), accompanied by a fairly high temperature (500° C.), is required to bring about this union, which takes place only in the presence of a catalyzer, such as specially prepared iron. The elements, before combination, are mixed in the exact molecular proportions required; that is, one volume of nitrogen to three volumes of hydrogen. Not over 8 per cent of the mixture at any one time combines to form ammonia. This is dissolved out with water. The uncombined gases are again sent through the process.

The Haber Process is notable as a successful method of forcing ordinarily inert nitrogen (obtained from liquid air) into combination.

The LeBlanc Process

This process, invented by LeBlanc at the time of the French Revolution, was the first successful method for producing sodium carbonate, Na_2CO_3, from common salt, $NaCl$.

Essentially, the process is as follows: Common salt is heated to a red heat with concentrated sulfuric acid, producing sodium sulfate (Glauber's Salt) and hydrogen chloride

(1) $$2\ NaCl + H_2SO_4 \longrightarrow 2\ HCl + Na_2SO_4.$$

The sodium sulfate is then mixed with coke and heated, when reduction of the sulfate takes place:

(2) $Na_2SO_4+2\,C \longrightarrow 2\,CO_2+Na_2S.$

The sodium sulfide is, in turn, treated at high temperature with limestone, $CaCO_3$:

(3) $Na_2S+CaCO_3 \longrightarrow CaS+Na_2CO_3.$

The resulting mass is mixed with water in which the sodium carbonate dissolves, after which it is separated from the solution by evaporation and crystallization.

The last two reactions usually take place at the same time, since in ordinary practice the sodium sulfate, coke, and limestone are heated together.

The hydrogen chloride, produced in the first reaction, is dissolved in water to produce hydrochloric acid, which is an exceedingly valuable by-product.

The LeBlanc Process is now almost entirely supplanted by the Solvay Process.

The Solvay Process

This process, now used almost exclusively in the manufacture of sodium carbonate, is economical because it makes nearly complete use of the by-products that are formed.

Carbon dioxide, CO_2, is allowed to bubble through a solution of common salt and ammonia, NH_3. Two reactions occur, one immediately after the other. In the first reaction ammonium bicarbonate, NH_4HCO_3, results.

(1) $NH_3+CO_2+H_2O \longrightarrow NH_4HCO_3.$

The first product, ammonium bicarbonate, immediately reacts with the salt to form sodium bicarbonate and ammonium chloride, NH_4Cl:

(2) $NH_4HCO_3+NaCl \longrightarrow NH_4Cl+NaHCO_3$.

The sodium bicarbonate, since it is only very slightly soluble, crystallizes out as a fine white solid.

The sodium bicarbonate is next decomposed by heat to form, as one of the products, sodium carbonate, Na_2CO_3:

(3) $2\ NaHCO_3 \longrightarrow H_2O+CO_2+Na_2CO_3$.

The ammonium chloride which is produced in operation (2) is treated with slaked lime, $Ca(OH)_2$, to produce ammonia, which thus becomes available for further use in operation (1):

(4) $Ca(OH)_2+2\ NH_4Cl \longrightarrow CaCl_2+2\ H_2O+2\ NH_3$.

The calcium hydroxide used to liberate the ammonia is obtained by treating quicklime, or calcium oxide, CaO, with water (slaking):

(5) $CaO+H_2O \longrightarrow Ca(OH)_2$.

The calcium oxide, needed for operation (5), is a product of the decomposition of limestone by heating:

(6) $CaCO_3 \longrightarrow CaO+CO_2$.

The carbon dioxide, one of the products of this last reaction, is used in operation (1).

Companies employing the Solvay Process are now strongly encouraging the use of calcium chloride, $CaCl_2$, the only by-product which is at present wasted to any extent.